

Chapter 6. Electronic Structure of Atoms

Media Resources

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Electromagnetic Spectrum
Bohr Model
Quantum Numbers
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Photoelectric Effect
Radial Electron Distribution
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Movies:

Flame Tests for Metals

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Other Resources

Further Readings:

Scientific American, **September 2004**
 Put Body to Them!
 Presenting the Bohr Atom
 Getting the Numbers Right—The Lonely Struggle
 of Rydberg
 Suitable Light Sources and Spectroscopes for
 Student Observation of Emission Spectra in
 Lecture Halls
 Niels Bohr
 100 Years of Quantum Mysteries
 On a Relation between the Heisenberg and
 deBroglie Principles
 Introducing the Uncertainty Principle Using
 Diffraction of Light Waves
 Perspectives on the Uncertainty Principle and
 Quantum Reality
 A Student's Travels, Close Dancing, Bathtubs,
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 Teaching Introductory Chemistry
 The Mole, the Periodic Table, and Quantum
 Numbers: An Introductory Trio
 Electron Densities: Pictorial Analogies for
 Apparent Ambiguities in Probability
 Calculations
 Mind over Matter
 Magnetic Whispers: Chemistry and Medicine
 Finally Tune Into Controversial Molecular
 Chatter
 Seeing Inside
 The Nobel Prize in Medicine for Magnetic
 Resonance Imaging
Chemistry in Britain, **June 1996**
 The Magnetic Eye
 Demystifying Introductory Chemistry Part 1:
 Electron Configurations from Experiment
 Quantum Analogies on Campus
 The Origin of the s, p, d, f Orbital Labels
 “New” Schemes for Applying the Aufbau Principle
 A Low-Cost Classroom Demonstration of the
 Aufbau Principle
 Housing Electrons: Relating Quantum Numbers,
 Energy Levels, and Electron Configurations
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 The Quantum Shoe Store and Electron Structure
 Some Analogies for Teaching Atomic Structure
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The Noble Gas Configuration—Not the Driving Force but the Rule of the Game in Chemistry
The Periodic Table as a Mnemonic Device for Writing Electronic Configurations
The Periodic Table and Electron Configurations

Live Demonstrations:

Simple and Inexpensive Classroom Demonstration of Nuclear Magnetic Resonance and Magnetic Resonance Imaging

6.8 Electron Configurations

6.9 Electron Configurations and the Periodic Table

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6.8 Electron Configurations

Chapter 6. Electronic Structure of Atoms

Common Student Misconceptions

- Some students have difficulty converting between angstroms, nanometers, etc. and meters.
- Students often have difficulty switching from the language of certainties to the language of probabilities.
- Students are often frightened or put off by the mathematics, vocabulary, foreign names, and an apparent intangibility of the information.
- Students are initially unaware that the quantum theory laid foundations for such areas as spectroscopy and nanotechnology, just to mention a few.
- Students confuse Bohr's *orbits* with *orbitals*; most spellcheckers do not recognize the word "orbital".
- Students mistakenly think that spectral lines represent energy levels; consequently...
- Students have difficulty associating a given line in an emission (or absorption) spectrum with a transition between *two* energy levels.
- When drawing the orbital diagrams, students often draw 2, 6, 10, and 14 "boxes" for *s*, *p*, *d*, and *f* orbitals, respectively.

Teaching Tips

- This is often students' first glimpse at the realm of quantum theory. They need to understand that the model has been built up to rationalize experimental data. They also need to know that elements of one theory are maintained in the subsequent theory.
- Students may not be familiar with the common symbol for wavelength (the lowercase lambda, λ) and the common symbol for frequency (the lowercase nu, ν).
- Using the unit s^{-1} for frequency makes the units cancel more easily.
- Many students will be familiar with the mnemonic for remembering the order of colors in the visible spectrum: ROY G BIV, which stands for red, orange, yellow, green, blue, indigo, and violet.
- A good analogy for the uncertainty principle: picture a busy intersection photographed at night. With a short exposure, you get a clear image of the position of every car, but you cannot tell how fast they are going or whether they are going forward or backward or about to swerve or turn. With a time-lapsed exposure, you can tell from the streaks of light the speed and direction of each car, but you cannot tell where each one currently is. You can know position or path, but not both.

Lecture Outline

6.1 The Wave Nature of Light^{1,2,3,4}

- The **electronic structure** of an atom refers to the arrangement of electrons.
- Visible light is a form of **electromagnetic radiation** or *radiant energy*.
- Radiation carries energy through space.
- Electromagnetic radiation is characterized by its wave nature.
- All waves have a characteristic **wavelength**, λ (lambda), and amplitude, *A*.
- The **frequency**, ν (nu), of a wave is the number of cycles that pass a point in one second.
 - The units of ν are *hertz* ($1 \text{ Hz} = 1 \text{ s}^{-1}$).

¹ Figure 6.3 from Transparency Pack

² Figure 6.4 from Transparency Pack

³ "Electromagnetic Spectrum" Activity from Instructor's Resource CD/DVD

⁴ Table 6.1 from Transparency Pack

- The speed of a wave is given by its frequency multiplied by its wavelength.
 - For light, speed, $c = \lambda \nu$,
 - Electromagnetic radiation moves through a vacuum with a speed of 3.00×10^8 m/s.
- Electromagnetic waves have characteristic wavelengths and frequencies.
- The *electromagnetic spectrum* is a display of the various types of electromagnetic radiation arranged in order of increasing wavelength.
 - Example: visible radiation has wavelengths between 400 nm (violet) and 750 nm (red).

FORWARD REFERENCES

- X-Ray diffraction will be discussed in Chapter 12.
- Light emitting diodes will be described in Chapter 11 (section 11.7).
- Different ranges of the electromagnetic spectrum will be mentioned in Chapters 18 and 23.
- Gamma radiation will be further discussed in Chapter 21.

6.2 Quantized Energy and Photons

- Some phenomena can't be explained using a wave model of light:
 - *Blackbody radiation* is the emission of light from hot objects.
 - The *photoelectric effect* is the emission of electrons from metal surfaces on which light shines.
 - *Emission spectra* are the emissions of light from electronically excited gas atoms.

Hot Objects and the Quantization of Energy

- Heated solids emit radiation (black body radiation)
 - The wavelength distribution depends on the temperature (i.e., “red hot” objects are cooler than “white hot” objects).
- Planck investigated black body radiation.
 - He proposed that energy can only be absorbed or released from atoms in certain amounts.
 - These amounts are called quanta.
 - A **quantum** is the smallest amount of energy that can be emitted or absorbed as electromagnetic radiation.
 - The relationship between energy and frequency is:

$$E = h \nu$$
 - where h is **Planck's constant** (6.626×10^{-34} J-s).
- To understand quantization consider the notes produced by a violin (continuous) and a piano (quantized):
 - A violin can produce any note when the fingers are placed at an appropriate spot on the bridge.
 - A piano can only produce notes corresponding to the keys on the keyboard.

The Photoelectric Effect and Photons^{5,6,7,8}

- The **photoelectric effect** provides evidence for the particle nature of light.
 - It also provides evidence for quantization.
- Einstein assumed that light traveled in energy packets called **photons**.
 - The energy of one photon is $E = h \nu$.
- Light shining on the surface of a metal can cause electrons to be ejected from the metal.
 - The electrons will only be ejected if the photons have sufficient energy (*work function*):
 - Below the threshold frequency no electrons are ejected.

⁵ **September 2004** issue of *Scientific American* from Further Readings

⁶ “Photoelectric Effect” Animation from Instructor’s Resource CD/DVD

⁷ Figure 6.7 from Transparency Pack

⁸ “Put Body to Them!” from Further Readings

- Above the threshold frequency, the excess energy appears as the kinetic energy of the ejected electrons.
- Light has wave-like AND particle-like properties.

FORWARD REFERENCES

- *Photoconductivity* in solar energy conversions and emission of photons by semiconductor nanoparticles will be described in Chapter 12 (section 12.9).
- *Photodissociation*, i.e. bond breaking as a result of an absorption of a photon by a molecule, as well as *photodecomposition* will be discussed in Chapter 18 (section 18.2).
- The role of photons from sunlight in *photosynthesis* will be further discussed in Chapter 23 (section 23.3).

6.3 Line Spectra and the Bohr Model

Line Spectra^{9,10,11}

- Radiation composed of only one wavelength is called *monochromatic*.
- Radiation that spans a whole array of different wavelengths is called *continuous*.
- When radiation from a light source, such as a lightbulb, is separated into its different wavelength components, a **spectrum** is produced.
 - White light can be separated into a **continuous spectrum** of colors.
 - A rainbow is a continuous spectrum of light produced by the dispersal of sunlight by raindrops or mist.
 - On the continuous spectrum there are no dark spots, which would correspond to different lines.
- Not all radiation is continuous.
 - A gas placed in a partially evacuated tube and subjected to a high voltage produces single colors of light.
 - The spectrum that we see contains radiation of only specific wavelengths; this is called a **line spectrum**.

Bohr's Model^{12,13}

- Rutherford assumed that electrons orbited the nucleus analogous to planets orbiting the sun.
 - However, a charged particle moving in a circular path should lose energy.
 - This means that the atom should be unstable according to Rutherford's theory.
- Bohr noted the line spectra of certain elements and assumed that electrons were confined to specific energy states. These were called orbits.
- Bohr's model is based on three postulates:
 - Only orbits of specific radii, corresponding to certain definite energies, are permitted for electrons in an atom.
 - An electron in a permitted orbit has a specific energy and is an "allowed" energy state.
 - Energy is only emitted or absorbed by an electron as it moves from one allowed energy state to another.
 - The energy is gained or lost as a photon.

⁹ Figure 6.9 from Transparency Pack

¹⁰ "Flame Tests for Metals" Movie from Instructor's Resource CD/DVD

¹¹ Figure 6.11 from Transparency Pack

¹² "Presenting the Bohr Atom" from Further Readings

¹³ "Bohr Model" Activity from Instructor's Resource CD/DVD

The Energy States of the Hydrogen Atom^{14,15,16}

- Colors from excited gases arise because electrons move between energy states in the atom.
- Since the energy states are quantized, the light emitted from excited atoms must be quantized and appear as line spectra.
- Bohr showed mathematically that

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

- where n is the *principal quantum number* (i.e., $n = 1, 2, 3, \dots \infty$) and R_{H} is the Rydberg constant.
- The product $hcR_{\text{H}} = 2.18 \times 10^{-18} \text{ J}$.
- The first orbit in the Bohr model has $n = 1$ and is closest to the nucleus.
- The furthest orbit in the Bohr model has $n = \infty$ and corresponds to $E = 0$.
- Electrons in the Bohr model can only move between orbits by absorbing and emitting energy in quanta ($E = h\nu$).
 - The **ground state** is the lowest energy state.
 - An electron in a higher energy state is said to be in an **excited state**.
- The amount of energy absorbed or emitted by moving between states is given by

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

Limitations of the Bohr Model^{17,18}

- The Bohr Model has several limitations:
 - It cannot explain the spectra of atoms other than hydrogen.
 - Electrons do not move about the nucleus in circular orbits.
- However, the model introduces two important ideas:
 - The energy of an electron is quantized: electrons exist only in certain energy levels described by quantum numbers.
 - Energy gain or loss is involved in moving an electron from one energy level to another.

FORWARD REFERENCES

- Absorption of sufficient amount of energy to ionize an atom will be further discussed in Chapter 7 (section 7.4).
- Emission of light with characteristic colors (flame test) by excited atoms of Li, Na, and K is shown in Chapter 7 (section 7.7).
- Selective absorption of light by chemicals, e.g., organic dyes, will be described in Chapter 9.
- Absorption of wavelengths from the visible part of the electromagnetic spectrum by molecules of chlorophyll and other pigments will be discussed in Chapter 23 (section 23.3).
- Absorption in the visible range will be responsible for colors of many transition metal complexes (Chapter 23, section 23.5).

¹⁴ “Getting the Numbers Right—The Lonely Struggle of Rydberg” from Further Readings

¹⁵ Figure 6.12 from Transparency Pack

¹⁶ “Suitable Light Sources and Spectroscopes for Student Observation of Emission Spectra in Lecture Halls” from Further Readings

¹⁷ “Niels Bohr” from Further Readings

¹⁸ “100 Years of Quantum Mysteries” from Further Readings

6.4 The Wave Behavior of Matter¹⁹

- Knowing that light has a particle nature, it seems reasonable to ask whether matter has a wave nature.
- This question was answered by Louis deBroglie.
- Using Einstein's and Planck's equations, deBroglie derived:

$$\lambda = h/mv$$

- The **momentum**, mv , is a particle property, whereas λ is a wave property.
 - **Matter waves** are the term used to describe wave characteristics of material particles.
 - Therefore, in one equation deBroglie summarized the concepts of waves and particles as they apply to low-mass, high-speed objects.
 - As a consequence of deBroglie's discovery, we now have techniques such as X-ray diffraction and electron microscopy to study small objects.

The Uncertainty Principle^{20,21}

- **Heisenberg's uncertainty principle**: we cannot determine the *exact* position, direction of motion, and speed of subatomic particles simultaneously.
- For electrons: we cannot determine their momentum and position simultaneously.

FORWARD REFERENCES

- X-ray diffraction and X-ray crystallography will be further discussed in Chapter 12 (section 12.2).

6.5 Quantum Mechanics and Atomic Orbitals^{22,23}

- Schrödinger proposed an equation containing both wave and particle terms.
- Solving the equation leads to **wave functions**, ψ .
- The wave function describes the electron's matter wave.
 - The square of the wave function, ψ^2 , gives the probability of finding the electron.
 - That is, ψ^2 gives the electron density for the atom.
 - ψ^2 is called the **probability density**.
- **Electron density** is another way of expressing probability.
 - A region of high electron density is one where there is a high probability of finding an electron.

Orbitals and Quantum Numbers^{24,25,26,27}

- If we solve the Schrödinger equation we get wave functions and energies for the wave functions.
- We call ψ **orbitals**.
- Schrödinger's equation requires three quantum numbers:
 - *Principal quantum number*, n . This is the same as Bohr's n .
 - As n becomes larger, the atom becomes larger and the electron is further from the nucleus.
 - *Angular momentum quantum number*, l . This quantum number depends on the value of n .

¹⁹ "On a Relation between the Heisenberg and deBroglie Principles" from Further Readings

²⁰ "Introducing the Uncertainty Principle Using Diffraction of Light Waves" from Further Readings

²¹ "Perspectives on the Uncertainty Principle and Quantum Reality" from Further Readings

²² Figure 6.16 from Transparency Pack

²³ "A Student's Travels, Close Dancing, Bathtubs, and the Shopping Mall: More Analogies in Teaching Introductory Chemistry" from Further Readings

²⁴ "The Mole, the Periodic Table and Quantum Numbers: An Introductory Trio" from Further Readings

²⁵ "Quantum Numbers" Activity from Instructor's Resource CD/DVD

²⁶ Table 6.2 from Transparency Pack

²⁷ Figure 6.17 from Transparency Pack

- The values of l begin at 0 and increase to $n - 1$.
- We usually use letters for l (s, p, d and f for $l = 0, 1, 2,$ and 3). Usually we refer to the s, p, d and f orbitals.
- This quantum number defines the shape of the orbital.
- *Magnetic quantum number, m_l .*
 - This quantum number depends on l .
 - The magnetic quantum number has integer values between $-l$ and $+l$.
 - There are $(2l+1)$ possible values of m_l .
 - For example, for $l = 1$, there are $(2 \times 1 + 1) = 3$ values of m_l : 0, +1, and -1.
 - Consequently, for $l = 1$, there are 3 orbitals: p_x, p_y and p_z .
 - Magnetic quantum numbers give the three-dimensional orientation of each orbital.
 - A collection of orbitals with the same value of n is called an **electron shell**.
 - There are n^2 orbitals in a shell described by a the n value.
 - For example, for $n = 3$, there are $3^2 = 9$ orbitals.
 - A set of orbitals with the same n and l is called a **subshell**.
 - Each subshell is designated by a number and a letter.
 - For example, $3p$ orbitals have $n = 3$ and $l = 1$.
 - There are n types of subshells in a shell described by a the n value.
 - For example, for $n = 3$, there are 3 subshells: $3s, 3p$ and $3d$.
 - Orbitals can be ranked in terms of energy to yield an Aufbau diagram.
 - Note that this Aufbau diagram is for a single electron system.
 - As n increases, note that the spacing between energy levels becomes smaller.

6.6 Representations of Orbitals^{28,29}

The s Orbitals^{30,31,32}

- All s orbitals are spherical.
- As n increases, the s orbitals get larger.
- As n increases, the number of **nodes** increases.
 - A node is a region in space where the probability of finding an electron is zero.
 - $\psi^2 = 0$ at a node.
 - For an s orbital the number of nodes is given by $n - 1$.
- We can plot a curve of *radial probability density* vs. distance (r) from the nucleus.
 - This curve is the **radial probability function** for the orbital.

The p Orbitals³³

- There are three p orbitals: p_x, p_y and p_z .
 - The three p orbitals lie along the $x-, y-,$ and $z-$ axes of a Cartesian system.
 - The letters correspond to allowed values of m_l of $-1, 0,$ and $+1$.
- The orbitals are dumbbell shaped; each has two *lobes*.
- As n increases, the p orbitals get larger.
- All p orbitals have a node at the nucleus.

²⁸ “The Origin of the s, p, d, f Orbital Labels” from Further Readings

²⁹ “Electron Densities: Pictorial Analogies for Apparent Ambiguities in Probability Calculations” from Further Readings

³⁰ “Radial Electron Distribution” Animation from Instructor’s Resource CD/DVD

³¹ Figure 6.18 from Transparency Pack

³² Figure 6.21 from Transparency Pack

³³ Figure 6.22 from Transparency Pack

The *d* and *f* Orbitals³⁴

- There are five *d* and seven *f* orbitals.
 - Three of the *d* orbitals lie in a plane bisecting the *x*-, *y*-, and *z*-axes.
 - Two of the *d* orbitals lie in a plane aligned along the *x*-, *y*-, and *z*-axes.
 - Four of the *d* orbitals have four lobes each.
 - One *d* orbital has two lobes and a collar.

FORWARD REFERENCES

- An overlap of atomic orbitals will be introduced in Chapter 9 (section 9.4).
- Hybridization of atomic orbitals will be discussed in Chapter 9 (section 9.5).
- Molecular orbitals will be introduced in Chapter 9 (section 9.7).
- Overlap of *p* orbitals on C atoms will be implicated in the formation of π bonds in organic chemistry, as mentioned in Chapters 9 and 22 (section 22.1; C vs. Si) and discussed in detail in Chapter 24.
- Energies of *d* orbitals in different crystal fields will be discussed in Chapter 23 (section 23.6).

6.7 Many-Electron Atoms

Orbitals and Their Energies^{35,36,37}

- In a many-electron atom, for a given value of *n*,
 - The energy of an orbital increases with increasing value of *l*.
- Orbitals of the same energy are said to be **degenerate**.

Electron Spin and the Pauli Exclusion Principle³⁸

- Line spectra of many electron atoms show each line as a closely spaced pair of lines.
- Stern and Gerlach designed an experiment to determine why.
 - A beam of atoms was passed through a slit and into a magnetic field and the atoms were then detected.
 - Two spots were found: one with the electrons spinning in one direction and one with the electrons spinning in the opposite direction.
- Since **electron spin** (electron as a tiny sphere spinning on its own axis) is quantized,
 - We define $m_s = \text{spin magnetic quantum number} = \pm 1/2$.
- **Pauli's exclusion principle** states that no two electrons can have the same set of 4 quantum numbers.
 - Therefore, two electrons in the same orbital must have opposite spins.

FORWARD REFERENCES

- The roles of screening and penetration in determining the relative energies of subshells within a shell will be explained in Chapter 7.
- Pauli's exclusion principle will also apply to hybrid orbitals in Chapter 9 (sections 9.4-9.5) and molecular orbitals (sections 9.7-9.8).

³⁴ Figure 6.23 from Transparency Pack

³⁵ Figure 6.24 from Transparency Pack

³⁶ “‘New’ Schemes for Applying the Aufbau Principle” from Further Readings

³⁷ “A Low-Cost Classroom Demonstration of the Aufbau Principle” from Further Readings

³⁸ “Mind over Matter” from Further Readings

6.8 Electron Configurations^{39,40,41,42,43}

- **Electron configurations** tell us how the electrons are distributed among the various orbitals of an atom.
- The most stable configuration, or ground state, is that in which the electrons are in the lowest possible energy state.
- When writing ground-state electronic configurations:
 - electrons fill orbitals in order of increasing energy with no more than two electrons per orbital.
 - no two electrons can fill one orbital with the same spin (Pauli).
 - for degenerate orbitals, electrons fill each orbital singly before any orbital gets a second electron.
 - How do we show spin?
 - An arrow pointing upwards has $m_s = + 1/2$ (spin up).
 - An arrow pointing downwards has $m_s = - 1/2$ (spin down).

Hund's Rule^{44,45,46,47,48,49,50,51,52,53}

- **Hund's rule:** for degenerate orbitals, the lowest energy is attained when the number of electrons with the same spin is maximized.
 - Thus, electrons fill each orbital singly with their spins parallel before any orbital gets a second electron.
 - By placing electrons in different orbitals, electron-electron repulsions are minimized.

Condensed Electron Configurations⁵⁴

- Electron configurations may be written using a shorthand notation (*condensed electron configuration*):
 - Write the **valence electrons** explicitly.
 - **Valence electrons** are electrons in the outer shell.
 - These electrons are gained and lost in reactions.
 - Write the **core electrons** corresponding to the filled noble gas in square brackets.
 - Core electrons are electrons in the inner shells.

³⁹ “Magnetic Whispers: Chemistry and Medicine Finally Tune Into Controversial Molecular Chatter” from Further Readings

⁴⁰ “Simple and Inexpensive Classroom Demonstration of Nuclear Magnetic Resonance and Magnetic Resonance Imaging” from Live Demonstrations

⁴¹ “Seeing Inside” from Further Readings

⁴² “The Nobel Prize in Medicine for Magnetic Resonance Imaging” from Further Readings

⁴³ **June 1996** issue of *Chemistry in Britain* from Further Readings

⁴⁴ “The Magnetic Eye” from Further Readings

⁴⁵ “Demystifying Introductory Chemistry; Part 1. Electron Configurations from Experiment” from Further Readings

⁴⁶ “Quantum Analogies on Campus” from Further Readings

⁴⁷ “Housing Electrons: Relating Quantum Numbers, Energy Levels, and Electron Configurations” from Further Readings

⁴⁸ “Pictorial Analogies VII: Quantum Numbers and Orbitals” from Further Readings

⁴⁹ “The Quantum Shoe Store and Electron Structure” from Further Readings

⁵⁰ Table 6.3 from Transparency Pack

⁵¹ “Electron Configuration” Activity from Instructor’s Resource CD/DVD

⁵² “Some Analogies for Teaching Atomic Structure at the High School Level” from Further Readings

⁵³ “Ionization Energies, Parallel Spins, and the Stability of Half-Filled Shells” from Further Readings

⁵⁴ “The Noble Gas Configuration—Not the Driving Force but the Rule of the Game in Chemistry” from Further Readings

- These are generally not involved in bonding.
- Example:
 - P is $1s^2 2s^2 2p^6 3s^2 3p^3$,
 - but Ne is $1s^2 2s^2 2p^6$.
 - Therefore, P is $[\text{Ne}]3s^2 3p^3$.

Transition Metals

- After Ar the d orbitals begin to fill.
- After the $3d$ orbitals are full the $4p$ orbitals begin to fill.
- The ten elements between Ti and Zn are called the **transition metals** or **transition elements**.

The Lanthanides and Actinides

- The $4f$ orbitals begin to fill with Ce.
 - Note: The electron configuration of La is $[\text{Xe}]6s^2 5d^1$.
- The $4f$ orbitals are filled for the elements Ce – Lu which are called **lanthanide elements** (or **rare earth elements**).
- The $5f$ orbitals are filled for the elements Th – Lr which are called **actinide elements**.
 - The actinide elements are radioactive and most are not found in nature.

FORWARD REFERENCES

- Periodic properties associated with electron configurations, such as atomic radii, ionization energies and electron affinities, will be discussed throughout Chapter 7.
- Valence electrons and the Octet Rule will be discussed in Chapter 8.
- Valence electrons of atoms within molecules and ions will be added and distributed according to the VSEPR model in Chapter 9 to determine molecular shapes.
- Electron configurations and the associated chemical properties of nonmetals in groups 4A–8A will be discussed in detail in Chapter 22.
- Electron configurations and the associated properties of select transition metals will be discussed in Chapter 23.
- High- and low-spin transition metal complexes will be discussed in Chapter 23 (section 23.6).
- Electron configuration of the C atom will be highlighted in Chapter 24 on organic chemistry.

6.9 Electron Configurations and the Periodic Table^{55,56,57,58,59}

- The periodic table can be used as a guide for electron configurations.
- The period number is the value of n .
- Groups 1A and 2A have their s orbitals being filled.
- Groups 3A–8A have their p orbitals being filled.
- The s -block and p -block of the periodic table contain the representative, or main-group, elements.
- Groups 3B–2B have their d orbitals being filled.
- The lanthanides and actinides have their f orbitals being filled.
 - The actinides and lanthanide elements are collectively referred to as the **f -block metals**.
- Note that the $3d$ orbitals fill after the $4s$ orbital. Similarly, the $4f$ orbitals fill after the $5d$ orbitals.
- In general, for representative elements we do not consider the electrons in completely filled d or f subshells to be valence electrons.

⁵⁵ Figure 6.30 from Transparency Pack

⁵⁶ “The Periodic Table as a Mnemonic Device for Writing Electronic Configurations” from Further Readings

⁵⁷ “The Periodic Table and Electron Configurations” from Further Readings

⁵⁸ Figure 6.31 from Transparency Pack

⁵⁹ “Periodic Table” Activity from Instructor’s Resource CD/DVD

- In general, for transition elements we do not consider the electrons in a completely filled f subshell to be valence electrons.

Anomalous Electron Configurations

- There are many elements that appear to violate the electron configuration guidelines.
 - Examples:
 - Chromium is $[\text{Ar}]3d^54s^1$ instead of $[\text{Ar}]3d^44s^2$.
 - Copper is $[\text{Ar}]3d^{10}4s^1$ instead of $[\text{Ar}]3d^94s^2$.
 - Half-full (d^5) and full (d^{10}) d subshells are particularly stable.

FORWARD REFERENCES

- Electron configurations of ions of the main group elements will be covered in Chapter 8.
- Electron configurations of transition metal cations will be mentioned in Chapter 8 and further used in Chapter 23.

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