

Electrons In Atoms Chapter 5 Test Review Answer

Name _____ Date _____ Class _____

5 ELECTRONS IN ATOMS

A. Matching
Match each item in Column A with the correct description in Column B. Write the letter of the correct item on the line.

Column A	Column B
_____ 1. the ways in which electrons are arranged around the nuclei of atoms	a. Aufbau principle
_____ 2. the ejection of electrons by metals when light shines on them	b. quantum
_____ 3. the region around the nucleus of an atom where an electron is likely to be located	c. photoelectric effect
_____ 4. An atomic orbital may describe at most two electrons.	d. atomic orbitals
_____ 5. the regions within which electrons have the highest probability of being found	e. Pauli exclusion principle
_____ 6. When electrons occupy orbitals of equal energy, one electron enters each orbital and all the orbitals contain one electron.	f. energy level
_____ 7. the amount of energy required to move an electron from its present energy level to the next highest one	g. Hund's rule
_____ 8. light quanta	h. electron configurations
_____ 9. Electrons enter orbitals of lowest energy first.	i. quantum mechanical model
_____ 10. the modern description of the electrons in atoms	j. photons

B. Multiple Choice
Choose the best answer and write its letter on the line.

_____ 11. Bohr's contribution to the development of atomic structure
a. was referred to as the "planetary model."
b. was the discovery that electrons surround a dense nucleus.
c. was proposed that electrons travel in circular orbits around the nucleus.
d. is the quantum mechanical model.

_____ 12. What is the total number of orbitals in the third principal energy level?
a. 1
b. 4
c. 9
d. 18

120 Core Teaching Resources

electrons in atoms chapter 5 test review answer

electrons in atoms chapter 5 test review answer guides you through the essential concepts of atomic structure, focusing on the behavior and properties of electrons. This comprehensive review aims to equip students with a solid understanding of electron configurations, quantum numbers, atomic orbitals, and the principles governing electron placement within atoms, as typically covered in a Chapter 5 chemistry or physics curriculum. We will delve into the quantum mechanical model, explaining the significance of each quantum number and their role in defining an electron's state. Furthermore, we'll explore the Aufbau principle, Hund's rule, and the Pauli exclusion principle, the fundamental rules that dictate how electrons fill atomic orbitals. Mastering these topics is crucial for understanding chemical bonding, molecular geometry, and the periodic trends of elements. Whether you're preparing for an upcoming exam or seeking to deepen your knowledge, this review provides clear explanations and answers to common questions related to electrons in atoms.

- Understanding the Quantum Mechanical Model of the Atom
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Unpacking the Quantum Mechanical Model of the

Atom

The modern understanding of electrons within atoms is best described by the quantum mechanical model. This model replaces the older Bohr model, which envisioned electrons orbiting the nucleus in fixed, planetary-like paths. Instead, the quantum mechanical model treats electrons as wave-like particles, their exact location and momentum cannot be known simultaneously. This principle is famously articulated by Heisenberg's Uncertainty Principle. The model utilizes mathematical functions called wave functions (ψ) to describe the probability of finding an electron in a particular region of space around the nucleus. These regions of high probability are known as atomic orbitals.

The quantum mechanical model provides a more accurate and detailed picture of atomic structure, explaining phenomena like spectral lines and chemical bonding. It acknowledges that electrons don't follow predictable orbits but rather exist in probability distributions. This probabilistic nature is a cornerstone of quantum mechanics and is fundamental to understanding the behavior of electrons in atoms. The energy levels of electrons are quantized, meaning they can only exist at specific discrete energy values, rather than a continuous range.

The Significance of Quantum Numbers

Quantum numbers are a set of numerical values that describe the properties of electrons in atoms. They are derived from the solutions to the Schrödinger equation and provide a unique "address" for each electron in an atom. Understanding these numbers is paramount to comprehending electron behavior and the arrangement of electrons within orbitals. Each electron in an atom is described by a unique set of four quantum numbers.

The Principal Quantum Number (n)

The principal quantum number, denoted by ' n ', is the most fundamental of the quantum numbers. It describes the electron's main energy level, also known as the electron shell. The value of ' n ' can be any positive integer, starting from 1 ($n=1, 2, 3, \dots$). As ' n ' increases, the electron's energy level increases, and the average distance of the electron from the nucleus also increases. For instance, electrons with $n=1$ are in the lowest energy level and closest to the nucleus, while electrons with $n=3$ are in a higher energy level and further from the nucleus.

The Azimuthal or Angular Momentum Quantum Number (l)

The azimuthal quantum number, denoted by ' l ', describes the shape of an atomic orbital and the subshell to which it belongs. For a given principal quantum number ' n ', the possible values of ' l ' range from 0 to $(n-1)$. The

values of ' l ' are commonly associated with letters: $l=0$ corresponds to an s orbital, $l=1$ to a p orbital, $l=2$ to a d orbital, and $l=3$ to an f orbital. So, for $n=1$, only $l=0$ is possible (1s orbital). For $n=2$, l can be 0 (2s orbital) or 1 (2p orbitals).

The Magnetic Quantum Number (m_l)

The magnetic quantum number, denoted by ' m_l ', specifies the orientation of an atomic orbital in three-dimensional space. For a given value of ' l ', the possible values of ' m_l ' range from $-l$ to $+l$, including 0. For example, if $l=0$ (s orbital), m_l can only be 0, indicating that there is only one s orbital. If $l=1$ (p orbitals), m_l can be -1 , 0, or $+1$, indicating three p orbitals oriented along the x, y, and z axes (p_x , p_y , p_z). If $l=2$ (d orbitals), m_l can be -2 , -1 , 0, $+1$, or $+2$, indicating five d orbitals with various spatial orientations.

The Spin Quantum Number (m_s)

The spin quantum number, denoted by ' m_s ', describes the intrinsic angular momentum of an electron, often visualized as the electron spinning on its axis. This spin generates a magnetic dipole moment. There are only two possible values for ' m_s ': $+1/2$ and $-1/2$, often represented as "spin up" and "spin down." This means that within a single atomic orbital, two electrons can exist, provided they have opposite spins. This fundamental concept is formalized by the Pauli Exclusion Principle.

Atomic Orbitals and Their Shapes

Atomic orbitals are regions of space around the nucleus where there is a high probability of finding an electron. Their shapes are determined by the azimuthal quantum number (l) and are often visualized as probability density plots. Each orbital can hold a maximum of two electrons, provided they have opposite spins.

The s Orbital

The s orbital corresponds to $l=0$ and is spherical in shape. Regardless of the principal quantum number ' n ', an s orbital is always centered on the nucleus. As ' n ' increases, the size of the s orbital increases, meaning the electron is found further from the nucleus on average. The 1s orbital is the smallest, followed by the 2s, 3s, and so on.

The p Orbitals

The p orbitals correspond to $l=1$ and have a dumbbell shape. For any given principal energy level ($n \geq 2$), there are three degenerate p orbitals (orbitals with the same energy). These are oriented along the x, y, and z axes, and are typically designated as p_x , p_y , and p_z . The nodal plane for each p orbital lies on the nucleus.

The d Orbitals

The d orbitals correspond to $l=2$ and have more complex shapes, typically cloverleaf-like, with four lobes and two nodal planes. For any given principal energy level ($n \geq 3$), there are five degenerate d orbitals. Four of these have four lobes, while the fifth, the d_{z^2} orbital, has a dumbbell shape with a torus (doughnut) around the middle.

The f Orbitals

The f orbitals correspond to $l=3$ and have even more intricate shapes, with multiple lobes and nodal planes. For any given principal energy level ($n \geq 4$), there are seven degenerate f orbitals. These are less commonly encountered in introductory chemistry but are important for understanding the properties of transition metals and lanthanides/actinides.

Electron Configuration and Filling Rules

Electron configuration describes the arrangement of electrons in atomic orbitals. This arrangement is governed by fundamental rules that ensure the lowest possible energy state for the atom. Understanding these rules is essential for predicting chemical behavior and drawing orbital diagrams.

The Aufbau Principle

The Aufbau principle, meaning "building up" in German, states that electrons fill atomic orbitals in order of increasing energy. Lower energy orbitals are filled before higher energy orbitals. This order is generally, but not always strictly, dictated by the principal quantum number ' n ' and the azimuthal quantum number ' l '. A common mnemonic for this order is the diagonal rule or Madelung rule, which shows the filling order as 1s, 2s, 2p, 3s, 3p, 4s, 3d, 4p, and so on.

Hund's Rule

Hund's rule states that within a subshell, electrons will occupy each orbital

singly with parallel spins before any orbital is doubly occupied. This minimizes electron-electron repulsion and leads to a more stable configuration. For example, when filling the three 2p orbitals, one electron will be placed in each orbital before a second electron is added to any of them. All these unpaired electrons will have the same spin.

The Pauli Exclusion Principle

The Pauli exclusion principle states that no two electrons in an atom can have the same set of four quantum numbers. This implies that an atomic orbital can hold a maximum of two electrons, and these two electrons must have opposite spins (one with $m_s = +1/2$ and the other with $m_s = -1/2$). If two electrons are in the same orbital, they must differ in their spin quantum number.

Electron Configurations of Common Elements

Applying the Aufbau principle, Hund's rule, and the Pauli exclusion principle allows us to determine the electron configurations for all the elements. This is often represented in a shorthand notation showing the occupied orbitals and the number of electrons in each. For instance, Helium (atomic number 2) has an electron configuration of $1s^2$. This means both its electrons are in the 1s orbital, with opposite spins.

Lithium (atomic number 3) has a configuration of $1s^2 2s^1$. The first two electrons fill the 1s orbital, and the third electron occupies the 2s orbital.

Carbon (atomic number 6) has a configuration of $1s^2 2s^2 2p^2$. According to Hund's rule, the two electrons in the 2p subshell will occupy two separate 2p orbitals with parallel spins.

Oxygen (atomic number 8) has a configuration of $1s^2 2s^2 2p^4$. The first two electrons fill the 2p subshell singly with parallel spins in two different orbitals, and the fourth electron pairs up with one of the existing electrons in a 2p orbital, with opposite spin.

Orbital Diagrams and Electron Spin

Orbital diagrams are a visual representation of electron configurations, using boxes or lines to represent atomic orbitals and arrows to represent electrons. Each box/line represents an orbital, and arrows pointing up or down represent electrons with their respective spins. For example, the orbital diagram for nitrogen (atomic number 7, $1s^2 2s^2 2p^3$) would show one electron in each of the three 2p orbitals, all with the same spin direction, demonstrating Hund's rule.

Understanding electron spin is crucial for correctly drawing orbital diagrams. A filled orbital will have one arrow pointing up and one arrow

pointing down within the same box. A partially filled orbital will have only one arrow, representing a single electron. The direction of the spin (up or down) for the first electron placed in a subshell is arbitrary, but subsequent electrons in that subshell, before pairing, must have the same spin according to Hund's rule.

Exceptions to Electron Configuration Rules

While the Aufbau principle and Hund's rule provide a general framework, some elements exhibit exceptions to these rules. These exceptions typically arise because half-filled or fully-filled subshells are more stable than partially filled ones. This extra stability can cause an electron to be promoted from a higher energy level to a lower energy subshell to achieve such a stable configuration.

A classic example is Chromium (atomic number 24). The predicted configuration based on the Aufbau principle is $[\text{Ar}] 4s^2 3d^4$. However, the actual configuration is $[\text{Ar}] 4s^1 3d^5$. This is because a half-filled d subshell ($3d^5$) is more stable than a partially filled one ($3d^4$). Similarly, Copper (atomic number 29) has an actual configuration of $[\text{Ar}] 4s^1 3d^{10}$, achieving a fully filled 3d subshell, instead of the predicted $[\text{Ar}] 4s^2 3d^9$.

These exceptions highlight the subtle interplay of energy levels and stability in determining electron arrangements. Students preparing for tests should be aware of these common exceptions, such as those found in groups 6 and 11 of the periodic table.

Relating Electron Configuration to Chemical Properties

The electron configuration of an atom is the primary determinant of its chemical behavior. The outermost electrons, known as valence electrons, are responsible for chemical bonding and reactivity. Elements with similar valence electron configurations tend to exhibit similar chemical properties, which is the basis of the periodic table's organization.

For example, elements in Group 1 (alkali metals) all have a single valence electron in their outermost s orbital (e.g., Lithium: $2s^1$, Sodium: $3s^1$). This single valence electron is easily lost, making these elements highly reactive metals that tend to form +1 ions.

Elements in Group 17 (halogens) all have seven valence electrons, typically with a configuration ending in $ns^2 np^5$ (e.g., Fluorine: $2s^2 2p^5$, Chlorine: $3s^2 3p^5$). They are one electron short of a stable octet, making them highly electronegative and prone to gaining an electron to form -1 ions. This understanding of how electron configurations translate to chemical properties is fundamental in chemistry.

Frequently Asked Questions

What does the atomic number of an element represent?

The atomic number represents the number of protons in the nucleus of an atom of that element. This number uniquely identifies an element.

Explain the concept of electron shells and subshells.

Electron shells (or energy levels) are regions around the nucleus where electrons are likely to be found, designated by principal quantum numbers ($n=1, 2, 3, \dots$). Subshells (s, p, d, f) are divisions within a shell, each having a specific shape and number of orbitals, holding a different maximum number of electrons.

What is the Aufbau principle, and how does it relate to electron configuration?

The Aufbau principle states that in the ground state of an atom, electrons fill atomic orbitals of the lowest available energy levels before occupying higher levels. This principle guides the filling order of electrons in subshells, determining the electron configuration of an atom.

Define Hund's rule and the Pauli exclusion principle.

Hund's rule states that within a subshell, electrons will singly occupy each orbital before pairing up. The Pauli exclusion principle states that no two electrons in an atom can have the same set of four quantum numbers, meaning an orbital can hold a maximum of two electrons with opposite spins.

What are valence electrons, and why are they important?

Valence electrons are the electrons in the outermost energy shell of an atom. They are important because they are involved in chemical bonding and determine the chemical properties of an element.

Describe the shape of s and p orbitals.

An s orbital is spherically symmetrical, meaning the probability of finding an electron in it is the same in all directions from the nucleus. A p orbital has a dumbbell shape, with two lobes on opposite sides of the nucleus along a specific axis (p_x, p_y, p_z).

What is ionization energy, and how does it trend across the periodic table?

Ionization energy is the minimum energy required to remove an electron from a gaseous atom or ion. It generally increases across a period from left to right (as nuclear charge increases and atomic radius decreases) and decreases down a group (as valence electrons are further from the nucleus and shielded by inner electrons).

Additional Resources

Here are 9 book titles related to "electrons in atoms chapter 5 test review answer," with short descriptions:

1. *The Quantum Leap: Understanding Electron Behavior*

This book delves into the fundamental principles governing the behavior of electrons within atomic structures. It breaks down complex quantum mechanical concepts, explaining how electrons occupy specific energy levels and orbitals. Readers will gain a clear understanding of the probabilistic nature of electron location and its implications for chemical bonding.

2. *Atomic Orbitals Explained: A Visual Guide*

Focusing on visual learning, this title provides detailed illustrations and diagrams of atomic orbitals. It clarifies the shapes and orientations of s, p, d, and f orbitals, making it easier to visualize electron distribution. The book connects these orbital concepts to the electron configurations of various elements.

3. *Electron Configurations Mastery: From Simple to Complex Atoms*

This comprehensive guide offers a systematic approach to determining electron configurations for all elements. It progresses from basic atomic models to the Aufbau principle, Hund's rule, and the Pauli exclusion principle. Practice problems and worked examples ensure a thorough grasp of how electrons fill atomic orbitals.

4. *The Periodic Table and Electron Shells*

This book explores the intricate relationship between the periodic table's structure and electron shell arrangements. It explains how the number of valence electrons and their orbital configurations dictate an element's chemical properties and placement. Understanding this connection is crucial for predicting reactivity and trends.

5. *Quantum Mechanics for Chemists: The Electron's Role*

A more in-depth exploration, this text bridges quantum mechanics and its application in chemistry. It rigorously explains the mathematical foundations behind atomic orbitals and the Schrödinger equation. The book highlights how quantum principles are essential for understanding molecular bonding and spectroscopic analysis.

6. *Valence Electrons and Chemical Bonding*

Dedicated to the outermost electrons, this book focuses on their critical role in forming chemical bonds. It details ionic, covalent, and metallic bonding, explaining how valence electrons are shared, transferred, or delocalized. The text emphasizes how electron behavior dictates molecular geometry and properties.

7. *Spectroscopy and Atomic Electron Transitions*

This title examines the interaction of light with atoms, specifically how it relates to electron energy levels. It explains the principles behind various spectroscopic techniques, like atomic emission and absorption spectroscopy. Understanding these transitions provides experimental evidence for electron configurations.

8. *The History of Atomic Theory: From Bohr to Schrödinger*

Tracing the evolution of our understanding of the atom, this book covers key historical developments. It highlights the contributions of scientists like Bohr, Rutherford, and Schrödinger in unraveling the nature of electrons. This historical perspective provides context for current atomic models and their limitations.

9. *Electron Affinity and Ionization Energy: Trends and Explanations*

This focused book examines two crucial periodic trends: electron affinity and ionization energy. It explains what these properties measure and how they are influenced by electron configuration and atomic structure. The text provides clear explanations for the observed trends across the periodic table.

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