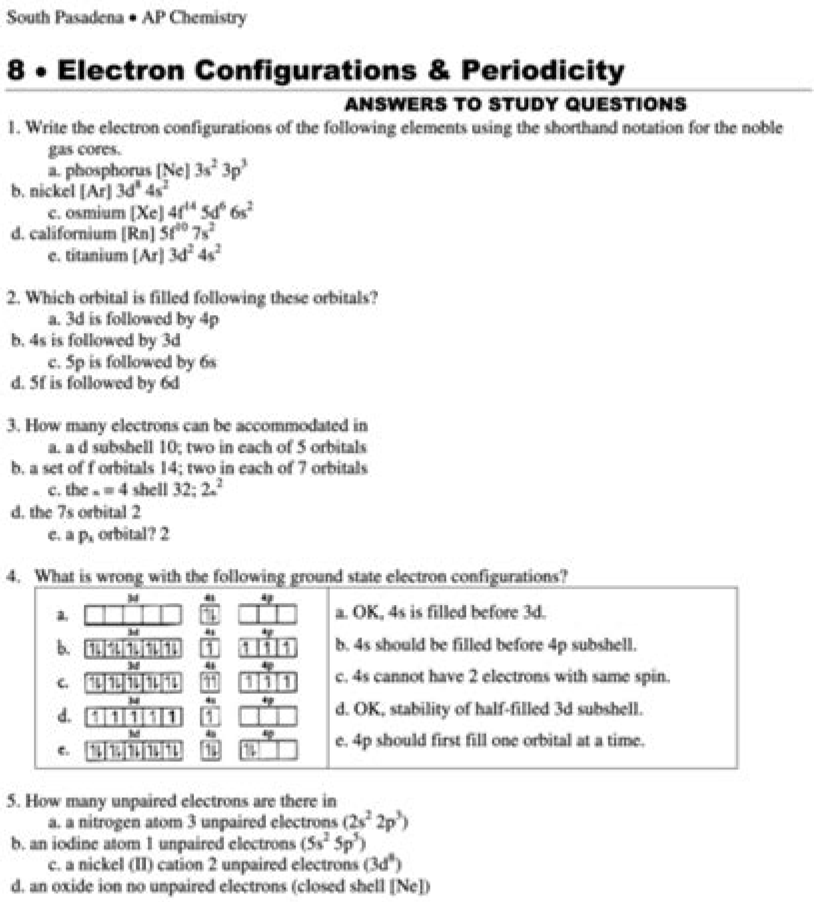
[Electrons And Periodicity Packet Answers Unit 3]



South Pasadena • AP Chemistry

8 • Electron Configurations & Periodicity

ANSWERS TO STUDY QUESTIONS

1. Write the electron configurations of the following elements using the shorthand notation for the noble gas cores.
a. phosphorus [Ne] $3s^2$ $3p^3$
b. nickel [Ar] $3d^8$ $4s^2$
c. osmium [Xe] $4f^{14}$ $5d^6$ $6s^2$
d. californium [Rn] $5f^{10}$ $7s^2$
e. titanium [Ar] $3d^2$ $4s^2$

2. Which orbital is filled following these orbitals?
a. 3d is followed by 4p
b. 4s is followed by 3d
c. 5p is followed by 6s
d. 5f is followed by 6d

3. How many electrons can be accommodated in
a. a d subshell 10; two in each of 5 orbitals
b. a set of f orbitals 14; two in each of 7 orbitals
c. the $n = 4$ shell 32; $2n^2$
d. the 7s orbital 2
e. a p_x orbital? 2

4. What is wrong with the following ground state electron configurations?

	3d	4s	4p	
a.		↑↓		a. OK, 4s is filled before 3d.
b.	↑↓ ↑↓ ↑↓ ↑↓ ↑↓	↑	↑ ↑ ↑	b. 4s should be filled before 4p subshell.
c.	↑↓ ↑↓ ↑↓ ↑↓ ↑↓	↑↑	↑ ↑ ↑	c. 4s cannot have 2 electrons with same spin.
d.	↑ ↑ ↑ ↑ ↑	↑		d. OK, stability of half-filled 3d subshell.
e.	↑↓ ↑↓ ↑↓ ↑↓ ↑↓	↑↓	↑↓	e. 4p should first fill one orbital at a time.

5. How many unpaired electrons are there in
a. a nitrogen atom 3 unpaired electrons ($2s^2$ $2p^3$)
b. an iodine atom 1 unpaired electrons ($5s^2$ $5p^5$)
c. a nickel (II) cation 2 unpaired electrons ($3d^8$)
d. an oxide ion no unpaired electrons (closed shell [Ne])

electrons and periodicity packet answers unit 3

electrons and periodicity packet answers unit 3 is a crucial resource for students and educators delving into the fundamental principles of atomic structure and the predictable patterns of elements. This comprehensive guide aims to provide detailed explanations and solutions to common queries found in such packets, focusing on electron configurations, quantum numbers, atomic radii, ionization energy, and electronegativity. Understanding these concepts is paramount for mastering chemistry, as they form the bedrock for predicting chemical behavior and bonding. This article will break down complex topics into digestible sections, offering clarity on trends within the periodic table and the underlying reasons for these patterns. We will explore the relationship between electron arrangement and the physical and chemical properties of elements, ensuring a thorough understanding of periodicity.

Understanding Electron Configurations and Quantum Numbers: Core Concepts for Periodicity

The Quantum Mechanical Model and Electron Orbitals

The foundation of understanding electron behavior and periodicity lies in the quantum mechanical model of the atom. Unlike the Bohr model, which depicted electrons in fixed orbits, the quantum mechanical model describes electrons in terms of probability distributions called orbitals. These orbitals are regions of space where there is a high probability of finding an electron. The shape and energy of these orbitals are defined by a set of quantum numbers.

Principal Quantum Number (n)

The principal quantum number, denoted by 'n', is the first of the quantum numbers and indicates the energy level of an electron. It can take on positive integer values (1, 2, 3, ...). Higher values of 'n' correspond to higher energy levels and electrons that are further from the nucleus. For example, electrons in the n=2 shell are in a higher energy level and generally further from the nucleus than those in the n=1 shell.

Angular Momentum Quantum Number (l)

The angular momentum quantum number, 'l', describes the shape of an electron's orbital and its subshell. The possible values of 'l' range from 0 to n-1. Different values of 'l' are associated with different orbital shapes:

- l = 0 corresponds to an s orbital, which is spherical in shape.
- l = 1 corresponds to a p orbital, which has a dumbbell shape.

- l = 2 corresponds to a d orbital, which has more complex shapes, often described as cloverleaf-like.
- l = 3 corresponds to an f orbital, which has even more intricate shapes.

Magnetic Quantum Number (ml)

The magnetic quantum number, 'ml', specifies the orientation of an orbital in space. For a given value of 'l', 'ml' can range from -l to +l, including 0. For example, if l = 1 (p orbitals), ml can be -1, 0, or +1, indicating the three different orientations of the p orbitals along the x, y, and z axes.

Spin Quantum Number (ms)

The spin quantum number, 'ms', describes the intrinsic angular momentum of an electron, often visualized as the electron spinning on its axis. Electrons can spin in one of two directions, represented by +1/2 or -1/2. This spin is crucial in determining the maximum number of electrons that can occupy a single orbital, as described by the Pauli Exclusion Principle.

Pauli Exclusion Principle and Electron Configurations

The Pauli Exclusion Principle states that no two electrons in an atom can have the same set of four quantum numbers. This means that an atomic orbital can hold a maximum of two electrons, and these two electrons must have opposite spins (one +1/2 and the other -1/2). Electron configurations systematically list the orbitals occupied by electrons in an atom, following specific rules like the Aufbau principle (filling lower energy orbitals first) and Hund's rule (maximizing unpaired electrons within a subshell).

Periodic Trends: Unpacking the Patterns of Electron Arrangement

Atomic Radius Trends

Atomic radius is a measure of the size of an atom, typically defined as half the distance between the nuclei of two bonded atoms of the same element. The periodic trends in atomic radius are directly related to electron configurations and the effective nuclear charge experienced by valence electrons.

Across a Period (Left to Right)

As you move from left to right across a period, the atomic radius generally decreases. This is because the number of protons in the nucleus increases, leading to a stronger positive charge. While electrons are added to the same principal energy level, the increased nuclear charge pulls the valence electrons more tightly towards the nucleus, resulting in a smaller atomic radius. The shielding effect from inner electrons remains relatively constant across a period.

Down a Group (Top to Bottom)

As you move down a group, the atomic radius generally increases. This is due to the addition of electrons to higher principal energy levels. These outer electrons are further from the nucleus and are shielded by the additional inner electron shells. Consequently, the attractive force of the nucleus on the outermost electrons is weaker, allowing the atomic radius to expand.

Ionization Energy Trends

Ionization energy is the minimum energy required to remove an electron from a gaseous atom or ion. The trends in ionization energy are largely inverse to those of atomic radius.

Across a Period (Left to Right)

Ionization energy generally increases across a period. This is because as the nuclear charge increases and the atomic radius decreases, the valence electrons are held more tightly by the nucleus. Therefore, more energy is needed to remove an electron. Elements with a stable electron configuration, like noble gases, tend to have very high ionization energies.

Down a Group (Top to Bottom)

Ionization energy generally decreases down a group. As the atomic radius increases, the outermost electrons are further from the nucleus and are better shielded by inner electrons. This weaker attraction means less energy is required to remove an electron. The ease of electron removal makes elements at the bottom of a group more reactive in losing electrons.

Electronegativity Trends

Electronegativity is a measure of the tendency of an atom to attract a bonding pair of electrons. It is a concept closely related to ionization energy but specifically applies to electrons involved in chemical bonds.

Across a Period (Left to Right)

Electronegativity generally increases across a period. Atoms with a greater effective nuclear charge attract bonding electrons more strongly. Nonmetals on the right side of the periodic table, with their strong pull on electrons, are highly electronegative, while metals on the left are less electronegative.

Down a Group (Top to Bottom)

Electronegativity generally decreases down a group. As the atomic radius increases and the valence electrons are further from the nucleus, the atom's ability to attract bonding electrons diminishes. The shielding effect of inner electrons also plays a significant role in reducing the effective nuclear charge experienced by valence electrons in larger atoms.

Electron Affinity Trends

Electron affinity is the energy change that occurs when an electron is added to a neutral atom in the gaseous state to form a negative ion. While there are some irregularities, general trends exist.

Across a Period (Left to Right)

Electron affinity generally becomes more negative (meaning energy is released) across a period. Atoms with a higher effective nuclear charge and smaller atomic radii have a greater attraction for an incoming electron. Halogens, for example, have highly negative electron affinities because they are one electron away from achieving a stable noble gas configuration.

Down a Group (Top to Bottom)

Electron affinity generally becomes less negative (or more positive, meaning energy is absorbed) down a group. The increasing atomic size and the shielding effect reduce the attraction for an incoming electron. The trend is not as consistent as others, with some exceptions, particularly between the second and third-period elements in certain groups.

Shielding Effect and Effective Nuclear Charge: Explaining the "Why" Behind Periodicity

The Shielding Effect Explained

The shielding effect, also known as electron shielding, refers to the reduction in the effective nuclear charge experienced by valence electrons due to the presence of inner-shell electrons. Inner electrons, being closer to the nucleus, are more effective at blocking or shielding the positive nuclear charge from the outer electrons. This shielding reduces the electrostatic attraction between the nucleus and the valence electrons.

Effective Nuclear Charge (Zeff)

Effective nuclear charge (Zeff) is the net positive charge experienced by an electron in an atom. It is calculated by subtracting the shielding effect (represented by a shielding constant, S) from the actual nuclear charge (Z, the atomic number): Zeff = Z - S. As you move across a period, the nuclear charge (Z) increases, but the shielding effect (S) from inner electrons increases only slightly. This results in a significant increase in Zeff. As you move down a group, the nuclear charge (Z) increases substantially, but the shielding effect (S) from the additional inner electron shells also increases substantially, keeping Zeff relatively constant or increasing slightly, but the increased distance from the nucleus dominates the trend.

Impact on Periodic Trends

The interplay between effective nuclear charge and the shielding effect is fundamental to understanding periodic trends. The increasing Zeff across a period leads to a stronger attraction for valence electrons, explaining the decrease in atomic radius and increase in ionization energy and electronegativity. Conversely, while Zeff may not change dramatically down a group, the increased distance of valence electrons from the nucleus and the enhanced shielding by more electron shells weaken the nuclear attraction, leading to larger atomic radii and lower ionization energies.

Electron Configurations and Orbitals: A Deeper Dive

Aufbau Principle, Hund's Rule, and the Pauli Exclusion Principle in Practice

These three principles are the guiding rules for filling atomic orbitals with electrons:

- The Aufbau Principle dictates that electrons fill atomic orbitals in order of increasing energy. Lower energy orbitals are filled before higher energy ones.

- Hund's Rule states that within a subshell, electrons will individually occupy each orbital before pairing up. Furthermore, each singly occupied orbital will have the same spin. This maximizes the total spin of the atom, contributing to stability.

- The Pauli Exclusion Principle, as mentioned earlier, ensures that no two electrons in an atom can have the same set of quantum numbers. This implies that each orbital can hold a maximum of two electrons, and these must have opposite spins.

Orbital Diagrams and Electron Filling

Orbital diagrams are visual representations of electron configurations, using boxes or lines to represent orbitals and arrows to represent electrons with their spins. For example, a p subshell with three orbitals (px, py, pz) would be depicted with three boxes. If there are three electrons in the p subshell, Hund's rule dictates that each box gets one arrow of the same spin. If there are four electrons, the fourth electron pairs up in one of the orbitals with an opposite spin.

Exceptions to Electron Configuration Rules

While the general rules are consistent, there are notable exceptions to the Aufbau principle, particularly for elements in the d-block and f-block. These exceptions often arise from the increased stability gained by achieving a half-filled or fully filled d or f subshell. For instance, chromium (Cr) has an electron configuration of [Ar] 4s1 3d5 instead of the predicted [Ar] 4s2 3d4. This is because a half-filled 3d subshell is more stable than a partially filled one. Similarly, copper (Cu) exhibits [Ar] 4s1 3d10, opting for a full 3d subshell over a partially filled one.

Solving Problems Related to "Electrons and Periodicity Packet Answers Unit 3"

Analyzing Electron Configurations of Ions

When forming positive ions (cations), electrons are removed from the outermost shell first, and typically from the highest principal quantum number (n). For example, to form Fe^{2+} from iron (Fe), which has the electron configuration [Ar] $4s^2$ $3d^6$, two electrons are removed from the 4s orbital, resulting in [Ar] $3d^6$. For negative ions (anions), electrons are added to the outermost unoccupied subshell according to the Aufbau principle.

Predicting Properties Based on Electron Configurations

An atom's position in the periodic table, determined by its electron configuration, directly dictates its chemical and physical properties. Elements with similar valence electron configurations tend to exhibit similar chemical behaviors. For example, alkali metals (Group 1) all have one valence electron in an s orbital (ns^1), making them highly reactive and prone to forming +1 ions. Understanding these links is key to answering many questions in an "electrons and periodicity packet."

Interpreting Quantum Numbers for Specific Elements

Given an element, one can determine the possible sets of quantum numbers for its valence electrons. For instance, for oxygen (atomic number 8), its electron configuration is $1s^2$ $2s^2$ $2p^4$. The valence electrons are in the n=2 shell. The 2p subshell has three orbitals (m_l = -1, 0, +1). According to Hund's rule, the four 2p electrons would occupy these orbitals such that there are two unpaired electrons and one paired set. Thus, a possible set of quantum numbers for one of the unpaired electrons could be n=2, l=1, m_l=0, m_s=+1/2.

Frequently Asked Questions

What is the relationship between electron configuration and an element's position on the periodic table?

An element's position on the periodic table is directly determined by its electron configuration. The period number corresponds to the highest energy level occupied by electrons, and the block (s, p, d, or f) indicates the subshell being filled. The group number often relates to the number of valence electrons.

How does atomic radius change across a period and down a group, and why?

Atomic radius generally decreases across a period from left to right due to increasing nuclear charge pulling valence electrons closer. It generally increases down a group because the addition of electron shells further shields the valence electrons from the nucleus, leading to a larger electron cloud.

Explain ionization energy and its periodic trends.

Ionization energy is the minimum energy required to remove an electron from a gaseous atom or ion. It generally increases across a period (as nuclear charge increases and atomic size decreases, making electrons harder to remove) and decreases down a group (as valence electrons are further from the nucleus and more shielded).

What is electron affinity, and how does it vary across the periodic table?

Electron affinity is the energy change that occurs when an electron is added to a gaseous atom. It generally becomes more negative (more energy released) across a period, as atoms with higher effective nuclear charge attract an additional electron more strongly. It's less predictable down a group but often becomes less negative or even positive.

Define electronegativity and describe its periodic trends.

Electronegativity is a measure of an atom's ability to attract electrons in a chemical bond. It increases across a period (due to stronger nuclear attraction for shared electrons) and decreases down a group (as valence electrons are further from the nucleus and more shielded).

How do valence electrons dictate an element's chemical behavior and group properties?

Valence electrons are the outermost electrons and are involved in chemical bonding. Elements in the same group have the same number of valence electrons, leading to similar chemical properties and reactivity patterns because they tend to form the same types of bonds and ions.

What are noble gases, and why are they largely unreactive?

Noble gases (Group 18) are largely unreactive because they have a full valence electron shell (octet rule), making them very stable and having high ionization energies and low electron affinities.

Explain the concept of effective nuclear charge and

its impact on atomic properties.

Effective nuclear charge (Zeff) is the net positive charge experienced by an electron in a multi-electron atom. It increases across a period as the nuclear charge increases and the shielding by inner electrons remains relatively constant. Higher Zeff leads to smaller atomic size and higher ionization energy.

How does electron shielding affect atomic size and ionization energy?

Electron shielding occurs when inner electrons repel outer electrons, reducing the effective nuclear charge experienced by the outer electrons. This shielding effect increases down a group, leading to larger atomic radii and lower ionization energies as valence electrons are further from the nucleus and less tightly held.

What are isoelectronic species, and how can their electron configurations be used to predict relative sizes?

Isoelectronic species are atoms or ions that have the same electron configuration. For isoelectronic species, the one with the greater nuclear charge (more protons) will have a stronger attraction for its electrons, resulting in a smaller ionic or atomic radius.

Additional Resources

Here are 9 book titles related to electrons and periodicity, with descriptions:

1. *Insights into Electron Configurations and the Periodic Table*
This book delves into the fundamental principles governing electron behavior within atoms. It clearly explains quantum numbers, orbital shapes, and how these dictate an element's position and properties on the periodic table. Readers will gain a deep understanding of the electron cloud model and its implications for chemical bonding.

2. *Illuminating Periodic Trends: From Atomic Radius to Electronegativity*
This title focuses on the predictable patterns observed across the periodic table. It meticulously details the reasons behind variations in atomic radius, ionization energy, electron affinity, and electronegativity. The text uses clear examples and visualizations to make these concepts accessible and memorable.

3. *Interpreting Spectroscopic Data and Electron Transitions*
This resource bridges the gap between theoretical electron behavior and experimental observation. It explores how atomic emission and absorption

spectra reveal electron energy levels and transitions. The book provides practical guidance on analyzing spectral data to identify elements and understand their electronic structures.

4. *Investigating Quantum Mechanical Models of the Atom*
This book offers a comprehensive exploration of the quantum mechanical model, moving beyond simpler Bohr models. It explains the Schrödinger equation and the wave functions that describe electron probability. The text discusses the development of atomic orbitals and their significance in predicting chemical properties.

5. *Illustrating Valence Electrons and Chemical Reactivity*
Focused on the outermost electrons, this book connects electron configuration to an atom's propensity to react. It explains how valence electrons determine the types of bonds an element will form and its general chemical behavior. The content is rich with examples illustrating reactivity patterns across different groups and periods.

6. *In-depth Analysis of Electron Shielding and Effective Nuclear Charge*
This title tackles the crucial concepts of electron shielding and effective nuclear charge. It explains how inner electrons reduce the attractive force experienced by outer electrons. This understanding is presented as key to interpreting many periodic trends.

7. *Intuitive Understanding of Electron Filling Order: Aufbau, Hund, and Pauli*
This book aims to demystify the rules that govern how electrons occupy atomic orbitals. It clearly explains the Aufbau principle, Hund's rule, and the Pauli exclusion principle with illustrative examples. The goal is to provide readers with an intuitive grasp of electron configurations.

8. *Implications of Electron Arrangement for Molecular Geometry*
This work extends the understanding of electron arrangements to the formation of molecules. It discusses how electron pairs (both bonding and non-bonding) around a central atom influence molecular shape. The book shows how valence electron configurations predict VSEPR theory outcomes.

9. *Introducing the s, p, d, and f Blocks: Electron Distribution Patterns*
This book categorizes elements based on the type of orbital their last valence electron enters. It provides a clear explanation of the s, p, d, and f blocks and their characteristic electron configurations. The text highlights how these blocks correspond to distinct chemical properties and behaviors.