

covalent bonding and polyatomic ions

Covalent Bonding and Polyatomic Ions: A Deep Dive into Chemical Connections

covalent bonding and polyatomic ions represent fundamental concepts in chemistry, crucial for understanding how atoms interact to form molecules and compounds. This article will demystify the intricate world of covalent bonds, where electrons are shared rather than transferred, and explore the unique nature of polyatomic ions, which are groups of atoms covalently bonded together but carrying an overall charge. We will delve into the principles governing covalent bond formation, including electronegativity and molecular geometry, and dissect the formation, naming, and reactivity of various polyatomic ions. Whether you're a student grappling with these concepts or a science enthusiast seeking a deeper understanding, this comprehensive guide will illuminate the beauty and complexity of these essential chemical building blocks.

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Understanding Covalent Bonding

Covalent bonding is the backbone of so much of the chemical world around us, from the water we drink to the air we breathe. At its core, covalent bonding involves the sharing of valence electrons between two or more atoms. This sharing allows each participating atom to achieve a more stable electron configuration, typically resembling that of a noble gas, which is characterized by a full outer electron shell. Unlike ionic bonds, where electrons are completely transferred from one atom to another, covalent bonds represent a collaborative effort, a chemical partnership where electrons are held in common by the nuclei of the bonded atoms.

This electron sharing creates a powerful attraction that holds the atoms together, forming discrete molecular units. The driving force behind covalent bond formation is the reduction of potential energy for the system. When atoms approach each other, their electron clouds begin to overlap. If this overlap leads to a lower overall energy state for the combined atoms, a stable covalent bond will form. The strength and characteristics of this bond are largely determined by the number of electron pairs shared and the nature of the atoms involved. We often visualize these shared electrons as occupying the space between the bonded nuclei.

Electronegativity and Polarity in Covalent Bonds

Electronegativity, a measure of an atom's ability to attract shared electrons in a chemical bond, plays a pivotal role in determining the nature of covalent bonds. When two atoms with identical or very similar electronegativities form a covalent bond, the electrons are shared equally. This results in a nonpolar covalent bond, where there is no significant separation of charge across the bond. A prime example of this is the bond between two identical atoms, like in the diatomic molecule O_2 .

However, when atoms with different electronegativities form a covalent bond, the electrons are not shared equally. The more electronegative atom pulls the shared electrons closer to its nucleus, creating a partial negative charge (δ^-) on that atom and a partial positive charge (δ^+) on the less electronegative atom. This unequal sharing leads to a polar covalent bond. Water (H_2O) is a classic illustration of this phenomenon; the oxygen atom is more electronegative than the hydrogen atoms, resulting in polar covalent bonds within the molecule. This polarity has profound implications for the physical and chemical properties of substances.

Molecular Geometry and VSEPR Theory

The arrangement of atoms in space, known as molecular geometry, is a direct consequence of covalent bonding and the repulsion between electron pairs. Valence Shell Electron Pair Repulsion (VSEPR) theory provides a powerful framework for predicting these shapes. According to VSEPR theory, electron pairs in the valence shell of a central atom, whether they are bonding pairs or lone pairs, repel each other and will arrange themselves as far apart as possible to minimize this repulsion. This fundamental principle dictates the bond angles and overall three-dimensional structure of a molecule.

For instance, a molecule with two electron domains around the central atom will adopt a linear geometry, while three electron domains will lead to a trigonal planar arrangement. Four electron domains typically result in a tetrahedral geometry. Understanding molecular geometry is crucial because it directly influences a molecule's polarity, its reactivity, and its ability to interact with other molecules. Even if a molecule contains polar covalent bonds, if its geometry is symmetrical, the bond dipoles can cancel each other out, making the entire molecule nonpolar. Conversely, an asymmetrical arrangement of polar bonds will result in a polar molecule.

Types of Covalent Bonds

Covalent bonds are not a monolithic entity; they come in different forms, each characterized by the number of electron pairs shared between atoms. These variations are fundamental to the diversity of chemical compounds we observe. Understanding these different types helps us predict bond strength, length, and reactivity, providing a deeper insight into molecular behavior.

Single, Double, and Triple Covalent Bonds

The simplest type of covalent bond is the single bond, formed when two atoms share one pair of electrons. This is the most common type of covalent bond and is represented by a single line between the atomic symbols, such as in the H-H bond of H_2 . Single bonds are relatively strong and flexible, allowing for rotation around the bond axis.

A double covalent bond occurs when two atoms share two pairs of electrons, represented by two parallel lines between the atomic symbols (e.g., $\text{O}=\text{O}$ in O_2). Double bonds are stronger and shorter than single bonds between the same two atoms. They also restrict rotation around the bond, leading to geometric isomerism in some cases. Finally, a triple covalent bond involves the sharing of three pairs of electrons between two atoms, indicated by three parallel lines (e.g., $\text{N}\equiv\text{N}$ in N_2). Triple bonds are the strongest and shortest of the three types and significantly limit rotational freedom.

Sigma (σ) and Pi (π) Bonds

Beyond the classification by the number of shared pairs, covalent bonds can also be categorized based on the region of overlap between atomic orbitals. Sigma (σ) bonds are formed by the direct, head-on overlap of atomic orbitals along the internuclear axis. This overlap is symmetrical around the bond axis and is characteristic of all single bonds. Sigma bonds are strong and allow for free rotation.

Pi (π) bonds, on the other hand, are formed by the sideways overlap of atomic orbitals, typically p orbitals, above and below the internuclear axis. This overlap is not symmetrical around the bond axis. Pi bonds are always present in double and triple bonds, in addition to the sigma bond. A double bond consists of one sigma bond and one pi bond, while a triple bond consists of one sigma bond and two pi bonds. The presence of pi bonds restricts rotation and contributes significantly to the overall strength and reactivity of the multiple bond.

Factors Influencing Covalent Bond Strength

The strength of a covalent bond isn't simply a matter of chance; it's governed by several key factors that dictate how tightly the atoms are held together. These factors are crucial for understanding chemical reactions, predicting stability, and designing new materials. When we talk about bond strength, we're essentially referring to the amount of energy required to break that bond.

Bond Length and Bond Energy

There's an inverse relationship between bond length and bond energy. Generally, shorter bonds are stronger bonds. Think of it like stretching a rubber band: the further you stretch it (longer it gets), the easier it is to break. Similarly, as atoms get closer together in a covalent bond, the attraction between their nuclei and the shared electrons becomes more intense, requiring more energy to pull them apart. For example, a carbon-carbon triple bond is shorter than a double bond, which is shorter than a

single bond, and consequently, the triple bond is also the strongest.

Bond energy, also known as bond dissociation enthalpy, is a quantitative measure of this strength. It's the energy required to break one mole of a particular type of bond in the gaseous state. Higher bond energy values indicate stronger, more stable bonds that are less likely to break under normal conditions. This concept is fundamental to thermochemistry and understanding the energy changes involved in chemical reactions.

Hybridization of Atomic Orbitals

In many molecules, the atomic orbitals involved in covalent bonding undergo a process called hybridization. Hybridization is the mixing of atomic orbitals to form new, equivalent hybrid orbitals that are oriented in specific directions to maximize orbital overlap and minimize electron repulsion. The type of hybridization adopted by an atom significantly influences the strength and geometry of the covalent bonds it forms.

For instance, sp^3 hybridized orbitals form stronger sigma bonds than unhybridized p orbitals because they have a greater s-character and are more directed towards the bonding partner. This is why carbon atoms forming four single bonds in methane (CH_4) result in a strong and stable molecule with a tetrahedral geometry. The nature of hybridization is directly linked to the number of sigma bonds and lone pairs around the central atom, providing a sophisticated model for understanding covalent bonding in a wide array of organic and inorganic compounds.

Introduction to Polyatomic Ions

While simple covalent bonds between two atoms are common, the chemical world is far richer, featuring entities known as polyatomic ions. These are not single atoms or simple diatomic molecules; instead, they are groups of two or more atoms covalently bonded together, but with an overall net electrical charge. This charge means they behave as a single unit in ionic compounds, influencing their solubility, reactivity, and overall properties.

Think of a polyatomic ion as a tightly knit family of atoms that carries a collective "credit" or "debt" in terms of electrons. This collective charge is the defining characteristic that distinguishes them from neutral molecules. Understanding polyatomic ions is essential because they are ubiquitous in many common substances, from household cleaning agents to biological molecules and geological formations. Their presence significantly expands the diversity of ionic compounds that can exist.

The Nature of Polyatomic Ions

At the heart of a polyatomic ion is a core of covalently bonded atoms. These atoms share electrons and hold together as a distinct structural unit. However, within this covalently bonded cluster, there's an imbalance of electrons relative to the total number of protons in the constituent atoms. This imbalance results in an overall charge for the entire group of atoms.

If the polyatomic ion has more electrons than protons, it carries a negative charge and is called an anion. If it has fewer electrons than protons, it carries a positive charge and is called a cation. These ions then interact with oppositely charged ions to form ionic compounds. For example, the sulfate ion (SO_4^{2-}) is a polyatomic anion, and it commonly combines with cations like sodium (Na^+) to form sodium sulfate (Na_2SO_4).

Distinguishing Polyatomic Ions from Molecules

It's important to draw a clear distinction between polyatomic ions and neutral polyatomic molecules. Both involve multiple atoms held together by covalent bonds. The key difference lies in their net charge. Neutral molecules, like water (H_2O) or carbon dioxide (CO_2), have an equal number of protons and electrons, resulting in an overall charge of zero. They exist as discrete, uncharged entities.

Polyatomic ions, conversely, have a net positive or negative charge. This charge dictates their behavior in chemical systems. They are attracted to oppositely charged ions and are key components of ionic compounds. For example, while ammonia (NH_3) is a neutral molecule, the ammonium ion (NH_4^+) is a polyatomic cation. This difference in charge fundamentally alters their chemical interactions and roles in compounds.

Formation and Stability of Polyatomic Ions

The formation of polyatomic ions involves intricate electron rearrangements, often stemming from reactions involving neutral molecules or the gain or loss of electrons by covalently bonded groups. Their stability is a complex interplay of bond strengths within the ion and the resonance structures that can be drawn to represent their electron distribution. Understanding these factors is key to predicting their existence and reactivity.

Electron Gain or Loss

Many polyatomic ions are formed through processes where a neutral molecule gains or loses electrons. For instance, oxygen can gain electrons to form the oxide ion (O^{2-}). Similarly, a neutral molecule like water can lose a proton (H^+) to form the hydroxide ion (OH^-). These processes involve significant energy considerations and are often facilitated by the surrounding chemical environment.

Consider the formation of the nitrate ion (NO_3^-) from nitric acid (HNO_3). While nitric acid is a neutral molecule, the dissociation of a proton leads to the formation of the nitrate ion, which carries a negative charge. The atoms within the nitrate ion – one nitrogen and three oxygens – remain covalently bonded, but the loss of the proton leaves the remaining structure with an excess of negative charge, distributed across the oxygen atoms.

The Concept of Resonance

Resonance is a critical concept for understanding the stability and bonding within many polyatomic ions. When a single Lewis structure cannot adequately represent the delocalization of electrons within a polyatomic ion, resonance structures are used. These are multiple valid Lewis structures that differ only in the placement of electrons, particularly in double bonds and lone pairs. The actual structure of the polyatomic ion is a hybrid average of all its contributing resonance structures.

This electron delocalization means that the bonds within the ion are often intermediate in length and strength between single and double bonds. For example, in the carbonate ion (CO_3^{2-}), resonance explains why all three carbon-oxygen bonds are identical in length and strength, even though a single Lewis structure might depict one double bond and two single bonds. This delocalization of charge and electron density contributes significantly to the stability of the polyatomic ion, making it less reactive than if the charge were localized on a single atom.

Common Polyatomic Ions and Their Properties

The world of chemistry is populated by a vast array of polyatomic ions, each with its unique characteristics and roles. While it's impossible to list them all, a focus on some of the most common ones provides a solid foundation for understanding their significance. These ions are the building blocks of many everyday substances, from baking soda to fertilizers.

Cations and Anions of Significance

Among the common polyatomic ions, ammonium (NH_4^+) stands out as a significant polyatomic cation. It is formed when ammonia (NH_3) accepts a proton (H^+). Ammonium ions are frequently found in fertilizers and cleaning products. On the anionic side, there's a wide variety. The hydroxide ion (OH^-) is fundamental to the concept of pH and is found in bases. The carbonate ion (CO_3^{2-}) is prevalent in rocks like limestone and is used in many industrial processes.

Other important polyatomic anions include nitrate (NO_3^-), a key component of fertilizers and explosives; sulfate (SO_4^{2-}), found in sulfuric acid and Epsom salts; and phosphate (PO_4^{3-}), essential for life as part of DNA and ATP. Halogen oxyanions, like chlorate (ClO_3^-) and perchlorate (ClO_4^-), are also notable for their oxidizing properties. Each of these ions, despite their shared feature of being covalently bonded groups with a charge, exhibits distinct chemical behaviors due to the specific atoms involved and their arrangement.

Applications and Occurrence

The practical applications and natural occurrences of polyatomic ions are extensive. Ammonium salts are widely used in agriculture as nitrogen fertilizers. Sulfates are found in detergents and in minerals like gypsum. Phosphates are crucial in biological systems for energy transfer and as structural

components of bones and teeth. Carbonates are the primary constituents of shells, corals, and many sedimentary rocks.

Nitrates are also important in biological processes, though excessive amounts can lead to environmental issues like eutrophication. The ability of these ions to form stable ionic compounds with various cations dictates their presence in diverse environments, from the oceans to the human body. Understanding their properties is therefore crucial not only for chemistry but also for fields like environmental science, biology, and geology.

Covalent Bonding Within Polyatomic Ions

The internal architecture of polyatomic ions is defined by robust covalent bonds. It's these intra-ion bonds that hold the group of atoms together as a cohesive unit, despite the presence of an overall charge. The strength and nature of these covalent bonds are what give polyatomic ions their distinct identities and stabilities.

Lewis Structures and Electron Counting

To understand the covalent bonding within a polyatomic ion, chemists often draw Lewis structures. This process involves counting the total number of valence electrons from all atoms in the ion, plus adding electrons for each negative charge or subtracting electrons for each positive charge. These electrons are then arranged as valence electrons and shared pairs to satisfy the octet rule (or duet rule for hydrogen) for as many atoms as possible.

For example, to draw the Lewis structure for the sulfate ion (SO_4^{2-}), we count 6 valence electrons for sulfur, $4 \times 6 = 24$ valence electrons for oxygen, and add 2 electrons for the 2- charge, totaling 32 valence electrons. By distributing these electrons as single bonds, double bonds, and lone pairs around the central sulfur atom, we can depict how the atoms are covalently linked. This exercise reveals the sharing of electrons that maintains the integrity of the polyatomic ion.

Bond Order and Delocalization

The concept of bond order, which represents the number of covalent bonds between two atoms, is particularly insightful when analyzing polyatomic ions that exhibit resonance. In resonance structures, a single bond in one representation might become a double bond in another, indicating that the actual bond order is an average. For instance, in the nitrate ion (NO_3^-), a single Lewis structure might show one $\text{N}=\text{O}$ double bond and two $\text{N}-\text{O}$ single bonds. However, through resonance, all three $\text{N}-\text{O}$ bonds are equivalent, with a bond order of approximately 1.33.

This delocalization of electron density over multiple atoms is a key contributor to the stability of polyatomic ions. It spreads the formal charges, reducing electrostatic repulsion within the ion and making it more resistant to dissociation or reaction. The more resonance structures an ion has, and the more symmetrical they are, the greater the degree of electron delocalization and, generally, the

greater the stability of the ion.

The Role of Polyatomic Ions in Compounds

Polyatomic ions are not isolated entities; they play a crucial role in the formation and properties of a vast number of chemical compounds. They act as the charged "building blocks" that combine with other ions to create electrically neutral ionic compounds. Their presence significantly expands the repertoire of substances that can exist and their associated characteristics.

Formation of Ionic Compounds

Ionic compounds are formed when positively charged ions (cations) and negatively charged ions (anions) are held together by electrostatic attraction. Polyatomic ions, whether cations like ammonium (NH_4^+) or anions like sulfate (SO_4^{2-}), frequently serve as one of these charged components. For example, sodium nitrate (NaNO_3) is an ionic compound formed from the sodium cation (Na^+) and the nitrate polyatomic anion (NO_3^-).

The formula of an ionic compound is determined by the charges of the ions involved, ensuring that the overall charge of the compound is zero. This often requires multiple polyatomic ions to balance the charge of another ion. For instance, aluminum sulfate [$\text{Al}_2(\text{SO}_4)_3$] involves two aluminum cations (each with a +3 charge) and three sulfate anions (each with a -2 charge), resulting in a net charge of $(2 \times +3) + (3 \times -2) = 0$.

Impact on Physical and Chemical Properties

The presence of polyatomic ions significantly influences the physical and chemical properties of ionic compounds. For instance, the relatively large size and complex structure of many polyatomic ions can affect the melting and boiling points of ionic solids. Compounds containing polyatomic ions often exhibit different solubilities in various solvents compared to those formed from simple monatomic ions. The specific arrangement of atoms within the polyatomic ion can also contribute to unique reactivity patterns.

Furthermore, the internal covalent bonds within a polyatomic ion are much stronger than the electrostatic attractions between ions in a crystal lattice. This means that polyatomic ions typically remain intact during chemical reactions, behaving as a single functional unit. For example, in a reaction involving sodium sulfate, the sulfate ion (SO_4^{2-}) is likely to remain as a sulfate ion, even as it forms new bonds with other species.

Predicting Formulas with Polyatomic Ions

Accurately predicting the chemical formula of compounds containing polyatomic ions is a fundamental skill in chemistry. It relies on a systematic approach that balances the charges of the constituent ions. This ability is essential for nomenclature, understanding reactions, and interpreting chemical data.

Charge Balancing Rules

The primary rule for predicting formulas with polyatomic ions is that the total positive charge from cations must equal the total negative charge from anions to form a neutral compound. This is achieved by determining the lowest common multiple of the charges. For instance, to combine aluminum ions (Al^{3+}) with sulfate ions (SO_4^{2-}), we need a total positive charge that is a multiple of 3 and a total negative charge that is a multiple of 2. The least common multiple of 3 and 2 is 6. Therefore, we need two aluminum ions ($2 \times +3 = +6$) and three sulfate ions ($3 \times -2 = -6$) to achieve electrical neutrality.

When a polyatomic ion is used more than once in a formula, it must be enclosed in parentheses, with the subscript outside the parentheses indicating the number of those ions. For example, the formula for aluminum sulfate is $\text{Al}_2(\text{SO}_4)_3$, not $\text{Al}_2\text{SO}_{43}$. This notation clearly distinguishes between the number of atoms of each element within the polyatomic ion and the number of polyatomic ions present in the compound.

Common Cation-Anion Pairings

Familiarity with common polyatomic ions and their charges is crucial for efficient formula prediction. Knowing that ammonium is NH_4^+ , hydroxide is OH^- , nitrate is NO_3^- , and so on, allows for rapid determination of compound formulas. For example, if you need to write the formula for potassium carbonate, you recall that potassium forms K^+ and carbonate is CO_3^{2-} . To balance the charges, you'll need two K^+ ions for every one CO_3^{2-} ion, resulting in the formula K_2CO_3 .

Similarly, understanding the charges of common cations, whether monatomic like magnesium (Mg^{2+}) or polyatomic like ammonium (NH_4^+), and common anions, like chloride (Cl^-) or phosphate (PO_4^{3-}), empowers you to construct correct chemical formulas. This systematic application of charge balancing and knowledge of common ion charges forms the bedrock of chemical nomenclature and formula writing.

Resonance in Polyatomic Ions

Resonance is a sophisticated concept that arises when a single Lewis structure fails to accurately depict the bonding within a molecule or ion. In the context of polyatomic ions, resonance is a vital tool for understanding electron distribution, bond strengths, and overall stability. It highlights the dynamic nature of electron sharing in these charged species.

Delocalized Electrons and Hybrid Structures

When a polyatomic ion exhibits resonance, its valence electrons, particularly those involved in pi bonding, are not confined to a single bond between two atoms. Instead, they are delocalized, meaning they are spread out over a larger area of the ion. This electron delocalization leads to a situation where the actual structure of the ion is an average or "hybrid" of multiple contributing Lewis structures, known as resonance structures. These resonance structures are theoretical representations, and the real ion is a composite of all of them.

For example, the carbonate ion (CO_3^{2-}) has three equivalent resonance structures. In each structure, there is one double bond and two single bonds between carbon and oxygen. However, experimentally, all three carbon-oxygen bonds are identical, with a length and strength intermediate between a typical single and double bond. This indicates that the electrons in the pi system are delocalized across all three oxygen atoms, creating a hybrid structure where the negative charge is distributed evenly among them. This delocalization significantly stabilizes the carbonate ion.

Impact on Bond Properties

The presence of resonance profoundly impacts the bond properties within polyatomic ions. As mentioned with the carbonate ion, resonance leads to bond lengths and bond energies that are averaged across the resonance structures. This means that bonds that might appear as single or double in individual Lewis structures are actually intermediate in character in the real polyatomic ion. This homogenization of bonds contributes to increased stability because the electron density is spread out, reducing localized areas of high electron concentration and thus minimizing repulsion.

The bond order, a measure of the number of electron pairs shared between two atoms, is also affected by resonance. In a molecule or ion with resonance, the bond order is not necessarily an integer. For the carbonate ion, each C-O bond has a bond order of approximately 1.33 (calculated as the total number of bonds in all resonance structures divided by the number of resonance structures). This intermediate bond order reflects the partial double-bond character distributed through resonance, leading to a more stable and uniformly bonded ion.

Understanding covalent bonding and polyatomic ions opens up a universe of chemical possibilities. From the simple sharing of electrons to the complex interplay of resonance within charged groups, these concepts are foundational to grasping the behavior of matter. The ability to predict how atoms combine, how ions form, and how they interact to create the substances around us is a testament to the elegant principles that govern chemistry. As we continue to explore the molecular world, the insights gained from studying covalent bonding and polyatomic ions will undoubtedly lead to further innovations and a deeper appreciation for the intricate dance of atoms.

Q: What is the difference between a covalent bond and an ionic bond?

A: The fundamental difference lies in how electrons are handled. In a covalent bond, electrons are shared between atoms to achieve stability. In an ionic bond, electrons are transferred from one atom

to another, creating oppositely charged ions that are then attracted to each other electrostatically.

Q: How do you determine if a bond is polar or nonpolar covalent?

A: The difference in electronegativity between the two bonded atoms determines polarity. If the electronegativity difference is very small or zero, the bond is nonpolar covalent (electrons shared equally). If there is a significant difference, the bond is polar covalent (electrons shared unequally), creating partial positive and negative charges.

Q: What is the octet rule, and how does it relate to covalent bonding?

A: The octet rule states that atoms tend to gain, lose, or share electrons to achieve a stable electron configuration with eight valence electrons, similar to noble gases. Covalent bonding achieves this stability by allowing atoms to share valence electrons, effectively filling their outer shells.

Q: Can a polyatomic ion contain only nonpolar covalent bonds?

A: Yes, it's possible, though less common for charged species. If all the atoms within the polyatomic ion have very similar electronegativities, the bonds between them could be nonpolar covalent. However, the overall charge would still mean it's an ion, not a neutral molecule.

Q: How are the charges on polyatomic ions determined?

A: The charge on a polyatomic ion is determined by the total number of valence electrons of its constituent atoms minus the total number of electrons needed for all atoms to achieve a stable electron configuration (usually an octet), and then accounting for any electron gain (negative charge) or loss (positive charge) that results in the ion.

Q: What is the significance of resonance in polyatomic ions?

A: Resonance indicates that electron density is delocalized over multiple atoms within the ion. This delocalization leads to increased stability, uniform bond lengths and strengths, and a more even distribution of charge, making the ion less reactive.

Q: How do you write the chemical formula for a compound containing a polyatomic ion?

A: You must balance the charges of the cation and anion. If a polyatomic ion needs to be used more than once to achieve neutrality, it is enclosed in parentheses with a subscript indicating the number of those ions.

Q: Are polyatomic ions always anions?

A: No, polyatomic ions can be either cations or anions. The ammonium ion (NH_4^+) is a common example of a polyatomic cation.

Q: What is the difference between a covalent bond within a polyatomic ion and the ionic bond holding the ion in a compound?

A: The covalent bonds are strong intramolecular forces that hold the atoms together within the polyatomic ion itself. The ionic bond is a weaker electrostatic attraction between the charged polyatomic ion and oppositely charged ions in a compound, holding the compound's crystal lattice together.

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