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I can explain the relationship between the type of bonding and the properties of the elements participating in the bond.Understand how electronegativity values of the representative elements increase going from left to right across a period and decrease going down a group. These trends can be understood qualitatively through the electronic structure of the atoms, the shell model, and Coulomb's law. Describe how valence electrons shared between atoms of similar electronegativity constitute a non-polar covalent bond. For example, bonds between carbon and hydrogen are effectively non-polar even though carbon is slightly more electronegative than hydrogen. Explain how valence electrons shared between atoms of unequal electronegativity constitute a polar covalent bond. a. The atom with a higher electronegativity will develop a partial negative charge relative to the other atom in the bond. b. In single bonds, greater differences in electronegativity lead to greater bond dipoles. c. All polar bonds have some ionic character, and the difference between ionic and covalent bonding is not distinct but rather a continuum.Explain how the difference in electronegativity is not the only factor in determining if a bond should be designated as ionic or covalent. Generally, bonds between a metal and nonmetal are ionic, and bonds between two nonmetals are covalent. Examination of the properties of a compound is the best way to characterize the type of bonding. Understand how in a metallic solid, the valence electrons from the metal atoms are considered to be delocalized and not associated with any individual atom. A chemical bond forms if the energy of combined atoms is lower than the energy of separated atoms. In general a bond will form when the two atoms bonded together will be lower in energy than the atoms separated. For example, the graph below shows how the potential energy changes when a bond is formed between two hydrogen atoms. It also shows that the potential energy of bonded hydrogen atoms is less than the potential energy of separate hydrogen atoms indicating that H2 molecule is more stable (has lower energy) than two separate H atoms. The potential energy start decreasing as the two atoms start interacting until a covalent bond is formed indicating the lowest overall energy of the system. There are three main types of Chemical bonds, ionic bonds, metallic bonds and covalent bonds. Ionic bonds: form between metals and nonmetals. Electrons are transferred from the metal to the non-metal. When a nonmetal and a metal react to form a binary ionic compound, the ions form so that the valence electron configuration of the nonmetal achieves the electron configuration of the next noble gas atom. The valence orbitals of the metal are emptied. Ionic bond is the attractive force between a positive ion and a negative ion. This type of attraction is a Coulombic or electrostatic attraction. Ionic bonds are stronger when the charges are larger and the ions are smaller, this is explained by Coulomb's Law., Properties of Ionic SubstancesForm crystals (lattice of positive and negative ions)High melting and boiling pointsHard, Brittle, Conduct electricity when dissolved and when molten (melted). Good insulators as a solidCovalent bonds are bonds that form between two nonmetals. In covalent bonds, electrons are shared between atoms. When electrons are equally shared between the bonded atoms a nonpolar covalent bond is formed. On the other side, polar covalent bonds form when electrons are not equally shared between the two bonded atoms. Polar covalent bonds are formed between two atoms of different electronegativity values. Because electrons are not equally shared in polar bonds the bond will have a dipole moment.Properties of Nonpolar and Polar Covalent Molecules (Non-metals)Non-lustrous,various colorsBrittle, hard or softPoor conductorsNonmetallic oxides are acidic and covalent Form anions by gaining electrons Metallic bonds form between metal atoms. It can be for just one type of metal, a pure substance, or for different types of metal, a mixture called an alloy. The metallic attractions are due to multiple metallic cations being attracted to a delocalized sea of valence electrons. The IMF is stronger when there are smaller metallic cations and when there are more valence elections. Properties of Metallic SubstancesShiny (Luster)Malleable and ductileConduct heat and electricityMetallic oxides are basic and ionicLose electrons to form cationsMetallic Bonds: In metallic bonds, electrons are mobile and they form a sea of electrons (This was covered in 1.7 Periodic Trends)Electronegativity is a measure of the ability of an atom (or group of atoms) to attract shared electrons. This can be explained by applying Coulomb's law.According to Coulomb's law, the attractive force between charged particles increases with an increase in charge and decreases with an increase in the distance.On the periodic table, electronegativity generally increases across a period and decreases down a group. Electronegativity decreases as you move down a column as there is a greater distance from the nucleus and because there is also more electron shielding. Electronegativity increases as you move across a period on the periodic table, from left to right. This is because the atomic radius is decreasing while the number of protons (and effective nuclear charge) is increasing.The range of electronegativity values is from 4.0 for fluorine (the most electronegative) to 0.7 for cesium (the least electronegative).Fluorine is the most electronegative element; it has a small radius with a small amount of electron shielding, coming from the 1s electrons,to the effective nuclear charge is high. This results in a strong pull on shared electronsTypes of Bonds and ElectronegativityWhether a bond will be ionic or covalent (polar or nonpolar covalent) is determined by how strongly the atoms involved attract shared electrons. What we call electronegativity. The greater the difference in electronegativity between two atoms them more ionic character the bond will have. We call bonds that are slightly ionic "polar covalent bonds". Polar covalent bonds mean that the atoms involved in the bond do not share the electrons equally and therefore have a positive end and a negative end or a positive pole and a negative pole.If lithium and fluorine react, which one has more attraction for an electron? Why?The fluorine has more attraction for an electron than does lithium. Both have valence electrons in the same principal energy level (the 2nd), but fluorine has a greater number of protons in the nucleus. Electrons are more attracted to a larger nucleus (if the principal energy level is the same).2. In a bond between fluorine and iodine, which has more attraction for an electron? Why?The fluorine also has more attraction for an electron than does iodine. In this case the nuclear charge of iodine is greater, but the valence electrons are at a much higher principal energy level (and the inner electrons shield the outer electrons).3. What is the general trend for electronegativity across rows and down columns on the periodic table?Generally speaking, the electronegativity increases in going from left to right across a period because the number of protons increases which increases the effective nuclear charge. The electronegativity decreases going down a group because the size of the atoms increases as you go down so when other electrons approach the larger atoms, the effective nuclear charge is not as great due to shielding from the current electrons present.The table below illustrates the relationship between electronegativity and bond type. 4. the following bonds from most to least polar: a) N–F O–F C–Fb) C–F N–O Si–Fc) Cl–Cl B–Cl S–Cl4. The greater the electronegativity difference between the atoms, the more polar the bond.a) C–F, N–F, O–Fb) Si–F, C–F, N–Oc) B–Cl, S–Cl, Cl–Cl5. Which of the following bonds would be the least polar yet still be considered polar covalent? Mg–O C–O O–O Si–O N–OThe correct answer is N-O. To be considered polar covalent, unequal sharing of electrons must still occur. Choose the bond with the least difference in electronegativity yet there is still some unequal sharing of electrons.6. Which of the following bonds would be the most polar without being considered ionic?Mg–O C–O O–O Si–O N–OThe correct answer is Si-O. To not be considered ionic, generally the bond needs to be between two nonmetals. The most polar bond between the nonmetals occurs with the bond that has the greatest difference in electronegativity.In polar bonds, electrons are not equally shared between atoms, creating a dipole moment. Dipole moment arises from differences in the electronegativities of bonded atoms. Molecules with dipole moment have a charge distribution that can be represented by a center of a positive charge and a center of a negative charge. An arrow is used to represent a dipole moment. The arrow points to the negative center with its tail starting from the positive center. Atoms form bonds to have a full valence shell of eight electrons (the octet rule). As atoms have a full valence shell, they become stable and have an electron configuration that resembles the nearest noble gas. When two nonmetals react to form a covalent bond, they share electrons in a way that completes the valence electron configurations of both atoms.When a nonmetal and a representative-group metal react to form a binary ionic compound, the ions form so that the valence electron configuration of the nonmetal achieves the electron configuration of the next noble gas atom. The valence orbitals of the metal are emptied.Please feel free to reach out to me with any questions or concerns. I look forward to an excellent year! Mrs. Amergamer@richmond.k12.mi.us ReviewsThis product has not yet been rated. Chemical bonds are the bonds between atoms that are formed during chemical reactions. These bonds allow substances containing two or more atoms to be formed. They are formed by electrostatic forces of attraction between the oppositely charged nuclei and electrons, or by a dipole attraction. Chemical bonds occur in the world all around us. Even the most basic things from oxygen to table salt require chemical bonds to form. There are many different types of chemical bonds, including covalent bonds, ionic bonds, metallic bonds and hydrogen bonds. Atoms form bonds in order to attain a stable electronic configuration. Oxygen (O2) is an example of a covalent bond, where atoms share their electrons. In contrast, table salt or sodium chloride (NaCl) is an example of an ionic bond, where an electron from an atom is transferred to another causing the two atoms to be positively and negatively charged, thus attracting each other. Welcome to the Chemistry Library. This Living Library is a principal hub of the LibreTexts project, which is a multi-institutional collaborative venture to develop the next generation of open-access texts to improve postsecondary education at all levels of higher learning. The LibreTexts approach is highly collaborative where an Open Access textbook environment is under constant revision by students, faculty, and outside experts to supplant conventional paper-based books. Campus BookshelvesBookshelvesLearning Objects Home is shared under a not declared license and was authored, remixed, and/or curated by LibreTexts. ▶Answer/ExplanationAnswer:(a) (i) One point earned for each Lewis electron-dot structureIndication of lone pairs of electrons are required on each structureResonance forms of CO_3^{2-} are not required(ii) In CO_2, the C-O interactions are double bonds, OR, in CO_3^{2-} the C-O interactions are resonance forms (or figures below.)The carbon-oxygen bond length is greater in the resonance forms than in the double bonds.1^{st}point earned for indicating double bonds are present in CO_2 OR resonance occurs in $CO_3^{2-}$$2^{nd}$ point earned for BOTH of the above AND indicating the relative lengths of the bond types(b) (i) One point earned for each Lewis electron-dot structureLone pairs of electrons are required on each structure(ii) CF_4 has a tetrahedral geometry, so the bond dipoles cancel, leading to a nonpolar molecule.With five pairs of electrons around the central S atom, SF_4 exhibits a trigonal bipyramidal electronic geometry, with the line pair of electrons. In this configuration, the bond dipoles do not cancel, and the molecules is polar.One point earnedfor each molecule for proper geometry and explanation When studying the types of chemical bonds for the AP Chemistry exam, you should focus on understanding the following concepts: the definition and formation of ionic, covalent, polar covalent, metallic, hydrogen bonds, and Van der Waals forces; the characteristics and properties of each bond type, including melting and boiling points, electrical conductivity, solubility, and physical properties; the electron transfer or sharing mechanisms that lead to bond formation; and how bond type influences the behavior and interactions of molecules. Additionally, you should be able to identify examples of compounds for each bond type, predict the type of bonding in a given substance based on its elements, and explain the significance of these bonds in real-world applications and biological systems. Chemical bonds are the attractive forces that hold atoms together to form molecules and compounds. These bonds are essential for the structure and stability of matter. Understanding the different types of chemical bonds—ionic, covalent, polar covalent, metallic, hydrogen bonds, and Van der Waals forces—is crucial for comprehending how substances interact and behave. Each bond type has unique characteristics and properties that influence the physical and chemical properties of compounds. Chemical bonds are the attractive forces that hold atoms or ions together in a molecule or compound. These bonds form as a result of the interactions between the electrons of the atoms involved, leading to the stability and formation of chemical substances. Chemical bonds can be classified into different types, such as ionic, covalent, polar covalent, metallic, hydrogen bonds, and Van der Waals forces, each with distinct properties and formation mechanisms. Definition: Ionic bonds form when one atom transfers electrons to another atom, resulting in the formation of positively and negatively charged ions. These opposite charges attract each other, creating a strong electrostatic force. Formation: Typically occurs between metals (which lose electrons) and non-metals (which gain electrons). Characteristics: Formation of Ions: Metal atoms lose electrons to form positive ions (cations), and non-metal atoms gain electrons to form negative ions (anions). Electronegativity Difference: Significant difference in electronegativity between the atoms involved. Crystal Lattice Structure: Ionic compounds form a regular repeating pattern of ions. Properties: High Melting and Boiling Points: Due to strong electrostatic forces between ions. Solubility in Water: Many ionic compounds dissolve in water. Electrical Conductivity: Conduct electricity when molten or dissolved in water due to the movement of ions. Examples: Sodium Chloride (NaCl): Na^+ and Cl^- ions. Magnesium Oxide (MgO): Mg^{2+} and O^{2-} ions. Calcium Fluoride (CaF_2): Ca^{2+} and F^- ions. Potassium Bromide (KBr): K^+ and Br^- ions. Aluminum Sulfide (Al_2S_3): Al^{3+} and S^{2-} ions. Definition: Covalent bonds form when two atoms share one or more pairs of electrons. This bond typically occurs between non-metal atoms with similar electronegativities. Formation: Between non-metals. Characteristics: Electron Sharing: Electrons are shared between atoms to achieve a stable electron configuration. Bond Strength: Can vary depending on the number of shared electron pairs (single, double, or triple bonds). Properties: Lower Melting and Boiling Points: Compared to ionic compounds. Variable Solubility: Solubility in water can vary; many covalent compounds are not soluble. Poor Electrical Conductivity: Do not conduct electricity in solid or liquid states. Examples: Water (H_2O): Two hydrogen atoms share electrons with one oxygen atom. Carbon Dioxide (CO_2): One carbon atom shares electrons with two oxygen atoms. Methane (CH_4): One carbon atom shares electrons with four hydrogen atoms. Oxygen (O_2): Two oxygen atoms share two pairs of electrons (double bond). Nitrogen (N_2): Two nitrogen atoms share three pairs of electrons (triple bond). Definition: A type of covalent bond where the electrons are shared unequally between the two atoms, resulting in a molecule with a partial positive charge on one end and a partial negative charge on the other. Formation: Between atoms with different electronegativities. Characteristics: Unequal Electron Sharing: Electrons are more attracted to the atom with higher electronegativity. Dipole Moment: Creates a partial charge difference within the molecule. Properties: Dipole-Dipole Interactions: Intermolecular forces between polar molecules. Solubility in Polar Solvents: Often soluble in water and other polar solvents. Higher Melting and Boiling Points: Compared to nonpolar covalent compounds. Examples: Water (H_2O): Oxygen is more electronegative than hydrogen, creating partial charges. Hydrogen Chloride (HCl): Chlorine is more electronegative than hydrogen. Ammonia (NH_3): Nitrogen is more electronegative than hydrogen. Sulfur Dioxide (SO_2): Oxygen is more electronegative than sulfur. Ethanol (C_2H_5OH): Oxygen is more electronegative than carbon and hydrogen. Definition: Metallic bonds form between metal atoms, where electrons are free to move throughout the structure, often described as a "sea of electrons." Formation: Between metal atoms. Characteristics: Delocalized Electrons: Electrons are not associated with any specific atom and can move freely. Conductivity: Metals are good conductors of electricity and heat due to free-moving electrons. Properties: Malleability and Ductility: Metals can be hammered into sheets and drawn into wires. Luster: Metals have a shiny appearance. High Melting and Boiling Points: Due to strong metallic bonding. Examples: Iron (Fe): Iron atoms share a pool of electrons. Copper (Cu): Copper atoms share a sea of electrons. Aluminum (Al): Aluminum atoms share a pool of electrons. Gold (Au): Gold atoms share a sea of electrons. Silver (Ag): Silver atoms share a pool of electrons. Definition: Hydrogen bonds are a type of weak intermolecular force that occurs between a hydrogen atom bonded to a highly electronegative atom (such as nitrogen, oxygen, or fluorine) and another electronegative atom. Formation: Between molecules, not within a molecule. Characteristics: Intermolecular Force: Weaker than ionic or covalent bonds but stronger than Van der Waals forces. Specific Conditions: Hydrogen must be bonded to N, O, or F. Properties: Higher Melting and Boiling Points: Compared to molecules without hydrogen bonds. Solubility in Water: Often soluble in water. Biological Importance: Stabilizes structures like DNA and proteins. Examples: Water (H_2O): Hydrogen bonds between hydrogen of one water molecule and oxygen of another. Ammonia (NH_3): Hydrogen bonds between hydrogen of one ammonia molecule and nitrogen of another. Hydrogen Fluoride (HF): Hydrogen bonds between hydrogen of one HF molecule and fluorine of another. Ethanol (C_2H_5OH): Hydrogen bonds between hydrogen of one ethanol molecule and oxygen of another. Proteins: Hydrogen bonds between amino acids stabilize protein structures. Definition: Van der Waals forces are weak intermolecular forces that include attractions between temporary dipoles in nonpolar molecules and permanent dipoles in polar molecules. Formation: Between all molecules, but more noticeable in nonpolar substances. Characteristics: Weak Interactions: Weaker than ionic, covalent, and hydrogen bonds. Types: Includes London dispersion forces (temporary dipoles) and dipole-dipole interactions (permanent dipoles). Properties: Low Melting and Boiling Points: Substances held together by Van der Waals forces have low melting and boiling points. Significance: Important in gases and nonpolar compounds. Examples: Helium (He): Weak forces between helium atoms in the gaseous state. Neon (Ne): Weak forces between neon atoms in the gaseous state. Methane (CH_4): Weak forces between nonpolar methane molecules. Carbon Tetrachloride (CCl_4): Weak forces between nonpolar CCl_4 molecules. Benzene (C_6H_6): Weak forces between benzene molecules. Lead Safety Advisor, Author, Writer, Speaker Bryan McWhorter is a safety professional with eight years of experience in driving and teaching safety. Bryan gained his knowledge and experience as... 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