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How to tell if a lewis structure has a dipole moment

These are problems using 3D molecules run in the application Jmol to help you visualize the molecule to determine if it is polar or non-polar. Step 1: Draw the Lewis structure. Note the number of electron regions around the central atom, and of these which are bonding or lone pairs (non-bonding pairs) Step 2: Use this info to determine the 3D geometry of the molecule. You do this by remembering "VSEPR". Valence Shell Electron Pair Repulsion theory: around the central atom all regions of electrons repel each other to get as far away from each other as possible while pivoting around the central atom. Note that double bonds and triple bonds count as a single region of electrons. Step 3: Determine if the molecular is polar or non-polar - a molecule is (i) non-polar if the charge distribution is symmetric and (ii) polar if the charge distribution is asymmetric (not symmetric). The Decision Process: After you draw the molecule in 3D representation using VSEPR rules, if the molecule has symmetry around the central atom, the bond dipole moments will "cancel out" (like pulling in opposite directions) and the molecule will therefore be non-polar. However, if the molecule is asymmetric, the bond dipole moments won't "cancel out" and the molecule will have a net dipole moment and the molecule is therefore polar. To really understand how to do this, the Lewis structure is only the first step. You need to consider the molecule in 3D (three dimensions). If you click on the example molecules (where it says 3D view) below you'll get a better understanding of why some molecules are polar and some not. Note: molecules with two atoms are not shown in these examples; they are always linear with sp hybridization. If the atoms are the same, the molecule is non-polar molecule; if the atoms are different, the molecule is polar. What good is this? The polarity of a molecule will tell you a lot about its solubility, boiling point, etc. when you compare it to other similar molecules. Water, for example, is a very light molecule (lighter than oxygen gas or nitrogen gas) and you might expect it would be a gas based on its molecular weight, however the polarity of water makes the molecules "stick together" very well. And it's a good thing, because if water was not so polar, we would certainly not be here. VSEPR Rules: Table of Molecular Geometry. Molecular Polarity Problems (with 3D solutions!). VSEPR Rules: Electron and Molecular Geometry On Central Atom. Click for Print View. Electron Regions, shape, & hybridization Bonding Regions Lone Pairs Electron Region Geometry Molecular Geometry Examples 2 sp 2 0 linear linear BeF2, CO2 1 1 linear CO, N2 ;N≡N: :C≡O: 3 sp2 3 0 trigonal planar trigonal planar BF3, CO32- 2 1 bent O3, SO2 1 2 linear O2 4 sp3 4 0 tetrahedral tetrahedral CH4, SO42- 3 1 trigonal pyramidal NH3, H3O+ 2 2 bent H2O, ClO2+ 1 3 linear HF, OH- 5 sp3d 5 0 trigonal bipyramidal trigonal bipyramidal PF5 4 1 seesaw SF4, TeCl4, IF4+ 3 2 T-shaped ClF3 2 3 linear I3-, XeF2 6 sp3d2 6 0 octahedral octahedral SF6, PF6-, SiF62- 5 1 square pyramidal BrF5, SbCl52- 4 2 square planar XeF4, ICl4- Remember!... Step 1: Draw the Lewis structure. Step 2: Draw the 3D molecular structure w/ VSEPR rules. Step 3: Use symmetry to determine if the molecule is polar or non-polar. Click on the molecule's name to see the answer, but first try to do it yourself! _____ NOTE: No Chem 1E03-01 students will have a formal tutorial during the week of November 2. You should try the problems on your own. Solutions will be posted. 1.a) Draw resonance forms for SO32- 26 valence electrons CO32- 24 valence electrons NO3- 24 valence electrons ClO3- 26 valence electrons ClO2- 20 valence electrons b) What are the bond orders for the central atom-to-oxygen bonds in these species? For SO32-, the bond order is 4/3. For CO32-, the bond order is 4/3. For NO3-, the bond order is 4/3. For ClO3-, the bond order is 5/3. For ClO2-, the bond order is 3/2.brs c) What is the average formal charge on the oxygen atoms in each of these species? The average formal charge on oxygen in SO32- is -2/3. The average formal charge on oxygen in CO32- is -2/3. The average formal charge on oxygen in NO3- is -2/3 (note that N has a formal charge of +1 in this ion). The average formal charge on oxygen in ClO3- is -1/3. The average formal charge on oxygen in ClO2- is -1/2. d) Choose all of the species of part a that violate the octet rule. The species that violate the octet rule are SO32-,ClO3-, and ClO2- e) Give the AXnEm classes and predict the shapes of all the anions. SO32- - AX3E - triangular pyramidal CO32- - AX3 - triangular planar NO3- - AX3 - triangular planar ClO3- - AX3E - triangular pyramidal ClO2- - AX2E2 - bent 2. Oxalic acid, (COOH)2, has two ionizable hydrogens and contains a carbon-carbon single bond. When one mole of oxalic acid reacts with one mole of calcium hydroxide,Ca(OH)2, the salt CaC2O4 is formed. a) Draw adequate Lewis structures to represent the bonding in the oxalate ion, C2O42-. There are 34 valence electrons. b) What is the carbon-oxygen bond order in oxalate ion? Bond order = (1 + 2 + 1 + 2)/4 = 1.5 c) What is the formal charge on each oxygen atom in oxalate ion? Formal charge on oxygen = (-1 + 0 + (-1) + 0)/4 = -1/2 3. N2O has a linear, unsymmetrical structure that may be thought of as a hybrid of two resonance forms. Draw the two resonance forms of N2O. There are 16 valence electrons. (Also remember that N and O obey the octet rule.) 4. A solid non-metallic element (A) from Group 5A was burned in excess oxygen to give a white solid (B) (Rxn 1). The white solid (B) dissolves in water to give a solution of a substance (C) (Rxn 2); (C) turns blue litmus red. To this solution of (C) was added an excess of aqueous KOH to produce an aqueous solution of (D) (Rxn 3). Write balanced equations for Rxns 1, 2, and 3; identify A, B, C, D. Reaction 1: P4(s) + 5O2(g) -> P4O10(s). Thus A is P4, and B is P4O10 Reaction 2:P4O10(s) + 6H2O(l) -> 4H3PO4(aq). Thus C is H3PO4. Reaction 3:H3PO4(aq) + 3KOH(aq) -> K3PO4(aq) + 3H2O(l). Thus D is K3PO4. 5. Phosphoric acid, H3PO4, contains three oxygen-hydrogen single bonds. When one mole of phosphoric acid reacts with one mole of sodium hydroxide, the salt sodium dihydrogen phosphate, NaH2PO4, results. a) Draw adequate Lewis structures for H3PO4 and H2PO4-. H3PO4 has 32 valence electrons. H2PO4- has 32 valence electrons. b) Calculate bond orders for all the phosphorus-oxygen bonds of H3PO4 and H2PO4-. In H3PO4, there is one P-O bond of bond order 2, and three P-O bonds of bond order 1. In H2PO4-, there are two P-O bonds of bond order 1.5, and two P-O bonds of bond order 1. c) Calculate the formal charges on each oxygen atom in H3PO4 and H2PO4-. In H3PO4, all oxygen atoms have a formal charge of zero. In H2PO4-, two oxygen atoms (bonded to H) have a formal charge of zero, while the other two have a formal charge of -1/2. 6. For the molecules OCS, ICl3, IF5, SO2, PF3, and CCl2Br2: a) determine the AXnEm class; b) decide which have dipole moments, and for those that do, decide the direction of the dipole moment. OCS has 16 valence electrons. Its Lewis structure (with bond moments shown) is It is an AX2 molecule, with a net dipole moment pointing to O. ICl3 has 28 valence electrons. Its Lewis structure (with bond moments shown) is It is an AX3E2 molecule, with a net dipole moment pointing from S to a point midway between the O atoms. PF3 has 26 valence electrons. Its Lewis structure is It is an AX3E molecule, with a net dipole moment pointing from P to the center of the triangle formed by the three fluorines. CCl2Br2 has 32 valence electrons. Its Lewis structure is It is an AX4 molecule, with a net dipole moment pointing from C and bisecting the Cl-C-Cl angle. 7. A gaseous molecule "X" is made by burning a yellow elemental solid in air. "X" is converted to "Y" by reaction with oxygen in the presence of a catalyst. "Y" reacts vigorously with water to produce "Z" which turns litmus red. a) Identify "X", "Y", and "Z". "X" = SO2, "Y" = SO3, and "Z" = H2SO4. b) Write a balanced equation for each reaction. S(s) + O2(g) -> SO2(g) 2SO2(g) + O2(g) -> 2SO3(g) SO3(g) + H2O(l) -> H2SO4(aq) c) Draw Lewis diagrams and indicate the shapes of the molecules "X" and "Y". SO2 has 18 valence electrons. Its Lewis structure is The molecule is AX2E, and is bent. SO3 has 24 valence electrons. Its Lewis structure is The molecule is AX3, and is triangular planar. d) State whether "X" and "Y" possess dipole moments and, if so, in which direction they point. SO2 has a dipole moment pointing from S to midway between the oxygen atoms. SO3 has no dipole moment. e) When 1 mole of "Z" is titrated with 2 moles of NaOH, a soluble salt "W" forms. Write an appropriate Lewis electron dot structure for the anion of this salt. "W" is SO42-. There are six resonance structures. One of them is Lone pairs (shown as pairs of dots) in the Lewis structure of hydroxide In chemistry, a lone pair refers to a pair of valence electrons that are not shared with another atom in a covalent bond[1] and is sometimes called an unshared pair or non-bonding pair. Lone pairs are found in the outermost electron shell of atoms. They can be identified by using a Lewis structure. Electron pairs are therefore considered lone pairs if two electrons are paired but are not used in chemical bonding. Thus, the number of lone pair electrons plus the number of bonding electrons equals the total number of valence electrons around an atom. Lone pair is a concept used in valence shell electron pair repulsion theory (VSEPR theory) which explains the shapes of molecules. They are also referred to in the chemistry of Lewis acids and bases. However, not all non-bonding pairs of electrons are considered by chemists to be lone pairs. Examples are the transition metals where the non-bonding pairs do not influence molecular geometry and are said to be stereochemically inactive. In molecular orbital theory (fully delocalized canonical orbitals or localized in some form), the concept of a lone pair is less distinct, as the correspondence between an orbital and components of a Lewis structure is often not straightforward. Nevertheless, occupied non-bonding orbitals (or orbitals of mostly nonbonding character) are frequently identified as lone pairs. Lone pairs in ammonia (A), water (B), and hydrogen chloride (C) A single lone pair can be found with atoms in the nitrogen group such as nitrogen in ammonia, two lone pairs can be found with atoms in the chalcogen group such as oxygen in water and the halogens can carry three lone pairs such as in hydrogen chloride. In VSEPR theory the electron pairs on the oxygen atom in water form the vertices of a tetrahedron with the lone pairs on two of the four vertices. The H-O-H bond angle is 104.5°, less than the 109° predicted for a tetrahedral angle, and this can be explained by a repulsive interaction between the lone pairs.[2][3][4] Various computational criteria for the presence of lone pairs have been proposed. While electron density ρ(r) itself generally does not provide useful guidance in this regard, the laplacian of the electron density is revealing, and one criterion for the location of the lone pair is where L(r) = -∇2ρ(r) is a local maximum. The minima of the electrostatic potential V(r) is another proposed criterion. Yet another considers the electron localization function (ELF).[5] Angle changes Tetrahedral Structure of Water The pairs often exhibit a negative polar character with their high charge density and are located closer to the atomic nucleus on average compared to the bonding pair of electrons. The presence of a lone pair decreases the bond angle between the bonding pair of electrons, due to their high electric charge which causes great repulsion between the electrons. They are also used in the formation of a dative bond. For example, the creation of the hydronium (H3O+) ion occurs when acids are dissolved in water and is due to the oxygen atom donating a lone pair to the hydrogen ion. This can be seen more clearly when looked at it in two more common molecules. For example, in carbon dioxide (CO2), the oxygen atoms are on opposite sides of the carbon, whereas in water (H2O) there is an angle between the hydrogen atoms of 104.5°. Due to the repulsive force of the oxygen atom's lone pairs, the hydrogens are pushed further away, to a point where the forces of all electrons on the hydrogen atom are in equilibrium. This is an illustration of the VSEPR theory. Dipole moments Lone pairs can make a contribution to a molecule's dipole moment. NH3 has a dipole moment of 1.47 D. As the electronegativity of nitrogen (3.04) is greater than that of hydrogen (2.2) the result is that the N-H bonds are polar with a net negative charge on the nitrogen atom and a smaller net positive charge on the hydrogen atoms. There is also a dipole associated with the lone pair and this reinforces the contribution made by the polar covalent N-H bonds to ammonia's dipole moment. In contrast to NH3, NF3 has a much lower dipole moment of 0.24 D. Fluorine is more electronegative than nitrogen and the polarity of the N-F bonds is opposite to that of the N-H bonds in ammonia, so that the dipole due to the lone pair opposes the N-F bond dipoles, resulting in a low molecular dipole moment.[6] Stereogenic lone pairs = Inversion a generic organic amine molecule at nitrogen A lone pair can contribute to the existence of chirality in a molecule, when three other groups attached to an atom all differ. The effect is seen in certain amines, phosphines,[7] sulfonium and oxonium ions, sulfoxides, and even carbanions. The resolution of enantiomers where the stereogenic center is an amine is usually precluded because the energy barrier for nitrogen inversion at the stereo center is low, which allow the two stereoisomers to rapidly interconvert at room temperature. As a result, such chiral amines cannot be resolved, unless the amine's groups are constrained in a cyclic structure (such as in Tröger's base). Unusual lone pairs A stereochemically active lone pair is also expected for divalent lead and tin ions due to their formal electronic configuration of ns2. In the solid state this results in the distorted metal coordination observed in the litharge structure adopted by both PbO and SnO. The formation of these heavy metal ns2 lone pairs which was previously attributed to intra-atomic hybridization of the metal s and p states[8] has recently been shown to have a strong anion dependence.[9] This dependence on the electronic states of the anion can explain why some divalent lead and tin materials such as PbS and SnTe show no stereochemical evidence of the lone pair and adopt the symmetric rocksalt crystal structure.[10][11] In molecular systems the lone pair can also result in a distortion in the coordination of ligands around the metal ion. The lead lone pair effect can be observed in supramolecular complexes of lead(II) nitrate, and in 2007 a study linked the lone pair to lead poisoning [12] Lead ions can replace the native metal ions in several key enzymes, such as zinc cations in the ALAD enzyme, which is also known as porphobilinogen synthase, and is important in the synthesis of heme, a key component of the oxygen-carrying molecule hemoglobin. This inhibition of heme synthesis appears to be the molecular basis of lead poisoning (also called "saturnism" or "plumbism").[13][14][15] Computational experiments reveal that although the coordination number does not change upon substitution in calcium-binding proteins, the introduction of lead disturbs the way the ligands organize themselves to accommodate such an emerging lone pair; consequently, these proteins are perturbed. This lone-pair effect becomes dramatic for zinc-binding proteins, such as the above-mentioned porphobilinogen synthase, as the natural substrate cannot bind anymore - in those cases the protein is inhibited. In Group 14 elements (the carbon group), lone pairs can manifest themselves by shortening or lengthening single (bond order 1) bond lengths.[16] as well as in the effective order of triple bonds as well.[17][18] The familiar alkynes have a carbon-carbon triple bond (bond order 3) and a linear geometry of 180° bond angles (figure A in reference [19]). However, further down in the group (silicon, germanium, and tin), formal triple bonds have an effective bond order 2 with one lone pair (figure B[19]) and trans-bent geometries. In lead, the effective bond order is reduced even further to a single bond, with two lone pairs for each lead atom (figure C[19]). In the organogermanium compound (Scheme 1 in the reference), the effective bond order is also 1, with complexation of the acidic isonitrile (or isocyanide) C-N groups, based on interaction with germanium's empty 4p orbital.[19][20] Different descriptions for multiple lone pairs Further information: Sigma-pi and equivalent-orbital models The symmetry-adapted and hybridized lone pairs of H2O In elementary chemistry courses, the lone pairs of water are described as "rabbit ears": two equivalent electron pairs of approximately sp3 hybridization, while the HOH bond angle is 104.5°, slightly smaller than the ideal tetrahedral angle of arccos(-1/3) ≈ 109.47°. The smaller bond angle is rationalized by VSEPR theory by ascribing a larger space requirement for the two identical lone pairs compared to the two bonding pairs. In more advanced courses, an alternative explanation for this phenomenon considers the greater stability of orbitals with excess s character using the theory of isovalent hybridization, in which bonds and lone pairs can be constructed with spx hybrids wherein nonintegral values of x are allowed, so long as the total amount of s and p character is conserved (one s and three p orbitals in the case of second-row p-block elements). To determine the hybridization of oxygen orbitals used to form the bonding pairs and lone pairs of water in this picture, we use the formula 1 + x cos θ = 0, which relates bond angle θ with the hybridization index x. According to this formula, the O-H bonds are considered to be constructed from O bonding orbitals of ≈sp4.0 hybridization (≈80% p character, ≈20% s character), which leaves behind O lone pairs orbitals of ≈sp2.3 hybridization (≈70% p character, ≈30% s character). These deviations from idealized sp3 hybridization for tetrahedral geometry are consistent with Bent's rule: lone pairs localize more electron density closer to the central atom compared to bonding pairs; hence, the use of orbitals with excess s character to form lone pairs (and, consequently, those with excess p character to form bonding pairs) is energetically favorable. However, theoreticians often prefer an alternative description of water that separates the lone pairs of water according to symmetry with respect to the molecular plane. In this model, there are two energetically and geometrically distinct lone pairs of water possessing different symmetry: one (o) in-plane and symmetric with respect to the molecular plane and the other (n) perpendicular and anti-symmetric with respect to the molecular plane. The o-symmetry lone pair (o(out)) is formed from a hybrid orbital that mixes 2s and 2p character, while the n-symmetry lone pair (p) is of exclusive 2p orbital parentage. The s character rich O (out) lone pair orbital (also noted nO(n)) is an ≈sp0.7 hybrid (≈40% p character, 60% s character), while the p lone pair orbital (also noted nO(n)) consists of 100% p character. Both models are of value and represent the same total electron density, with the orbitals related by a unitary transformation. In this case, we can construct the two equivalent lone pair hybrid orbitals h and h' by taking linear combinations h = c1o(out) + c2p and h' = c1o(out) - c2p for an appropriate choice of coefficients c1 and c2. For chemical and physical properties of water that depend on the overall electron distribution of the molecule, the use of h and h' is just as valid as the use of o(out) and p. In some cases, such a view is intuitively useful. For example, the stereoelectronic requirement for the anomeric effect can be rationalized using equivalent lone pairs, since it is the overall donation of electron density into the antibonding orbital that matters. An alternative treatment using o/n separated lone pairs is also valid, but it requires striking a balance between maximizing nO(n)-σ* overlap (maximum at 90° dihedral angle) and nO(o)-σ* overlap (maximum at 0° dihedral angle), a compromise that leads to the conclusion that a gauche conformation (60° dihedral angle) is most favorable, the same conclusion that the equivalent lone pairs model rationalizes in a much more straightforward manner.[21] Similarly, the hydrogen bonds of water form along the directions of the "rabbit ears" lone pairs, as a reflection of the increased availability of electrons in these regions. This view is supported computationally.[5] However, because only the symmetry-adapted canonical orbitals have physically meaningful energies, phenomena that have to do with the energies of individual orbitals, such as photochemical reactivity or photoelectron spectroscopy, are most readily explained using σ and n lone pairs that respect the molecular symmetry.[21][22] Because of the popularity of VSEPR theory, the treatment of the water lone pairs as equivalent is prevalent in introductory chemistry courses, and many practicing chemists continue to regard it as a useful model. A similar situation arises when describing the two lone pairs on the carbonyl oxygen of a ketone.[23] However, the question of whether it is conceptually useful to derive equivalent orbitals from symmetry-adapted ones, from the standpoint of bonding theory and pedagogy, is still a controversial one, with recent (2014 and 2015) articles opposing[24] and supporting[25] the practice. See also Wikimedia Commons has media related to Lone pair. Coordination complex Highest occupied molecular orbital Inert pair effect Ligand Shared pair References ↑ IUPAC Gold Book definition: lone (electron) pair ↑ Fox, M.A.; Whitesell, J.K. (2004). Organic Chemistry. Jones and Bartlett Publishers. ISBN 978-0-7637-2197-8. Retrieved 5 May 2021. ↑ McMurry, J. (2000). Organic Chemistry 5th Ed. Cengage Learning India Pvt Limited. ISBN 978-81-315-0039-2. Retrieved 5 May 2021. ↑ Lee, J. D. (1968). Concise Inorganic Chemistry. Student's paperback edition. Van Nostrand. Retrieved 5 May 2021. ↑ a b Kumar, Anmol; Gadre, Shridhar R.; Mohan, Neetha; Suresh, Cherumuttathu H. (2014-01-06). "Lone Pairs: An Electrostatic Viewpoint". The Journal of Physical Chemistry A. 118 (2): 526–532. Bibcode:2014JPCA..118..526K. doi:10.1021/jp4117003. ISSN 1089-5639. PMID 24372481. ↑ Housecroft, C. E.; Sharpe, A. G. (2004). Inorganic Chemistry (2nd ed.). Prentice Hall, p. 40. ISBN 978-0-13-039913-7. ↑ Quin, L. D. (2000). A Guide to Organophosphorus Chemistry. LOCATION: John Wiley & Sons. ISBN 0471318248. ↑ Stereochemistry of Ionic Solids,J.D.Dunitz and L.E.Orgel, Advan. Inorg. and Radiochem. 1960, 2, 1-60 ↑ Payne, D. J. (2006). "Electronic Origins of Structural Distortions in Post-Transition Metal Oxides: Experimental and Theoretical Evidence for a Reversion of the Lone Pair Model". Physical Review Letters. 96 (15): 157403. doi:10.1103/PhysRevLett.96.157403. PMID 16712195. ↑ Walsh, Aron (2005). "The origin of the stereochemically active Pb(II) lone pair: DFT calculations on PbO and PbS". Journal of Solid State Chemistry. 178 (5): 1422–1428. Bibcode:2005JSSCh.178.1422W. doi:10.1016/j.jssc.2005.01.030. ↑ Walsh, Aron (2005). "Influence of the Anion on Lone Pair Formation in Sn(II) Monochalcogenides: A DFT Study". The Journal of Physical Chemistry B. 109 (40): 18868–18875. doi:10.1021/jp051822z. PMID 16853428. ↑ Gourlaouen, Christophe; Parisel, Olivier (15 January 2007). "Is an Electronic Shield at the Molecular Origin of Lead Poisoning? A Computational Modeling Experiment". Angewandte Chemie International Edition. 46 (4): 553–556. doi:10.1002/anie.200603037. PMID 17152108. ↑ Jaffe, E. K.; Martins, J.; et al. (13 October 2000). "The Molecular Mechanism of Lead Inhibition of Human Porphobilinogen Synthase". Journal of Biological Chemistry. 276 (2): 1531–1537. doi:10.1074/jbc.M007663200. PMID 11032836. ↑ Scinciarrello, Franco; Murray, H. Edward; et al. (15 September 2006). "Lead and 6-Aminolevulinic Acid Dehydratase Polymorphism: Where Does it Lead? A Meta-Analysis". Environmental Health Perspectives. 115 (1): 35–41. doi:10.1289/ehp.9448. PMC 1797830. PMID 17366816. ↑ Chhabra, Namrata (November 15, 2015). "Effect of Lead poisoning on heme biosynthetic pathway". Clinical Cases: Biochemistry For Medics. Archived from the original on 3 April 2016. Retrieved 30 October 2016. ↑ Richards, Anne F.; Brynda, Marcin; Power, Philip P. (2004). "Effects of the alkali metal counter ions on the germanium-germanium double bond length in a heavier group 14 element ethenide salt". Chem. Commun. (14): 1592–1593. doi:10.1039/B401507J. PMID 15263933. ↑ Power, Philip P. (December 1999). "π-Bonding and the Lone Pair Effect in Multiple Bonds between Heavier Main Group Elements". Chemical Reviews. 99 (12): 3463–3504. doi:10.1021/cr9408989. PMID 11849028. ↑ Vladimir Ya. Lee; Akira Sekiguchi (22 July 2011). Organometallic Compounds of Low-Coordinate Si, Ge, Sn and Pb: From Phantom Species to Stable Compounds. John Wiley & Sons, p. 23. ISBN 978-1-119-95626-6. ↑ a b c d Spikes, Geoffrey H.; Power, Philip P. (2007). "Lewis base induced tuning of the Ge-Ge bond order in a "digermene"". Chem. Commun. (1): 85–87. doi:10.1039/b612202g. PMID 17279269. ↑ Power, Philip P. (2003). "Silicon, germanium, tin and lead analogues of acetylenes". Chemical Communications (17): 2091–101. doi:10.1039/B212224C. PMID 13678155. ↑ a b A. Albricht, Thomas (2013-04-08). Orbital interactions in chemistry. Burdett, Jeremy K., 1947-, Whangbo, Myung-Hwan (Second ed.). Hoboken, New Jersey. ISBN 9780471080398. OCLC 823294395. ↑ While nO(n) lone pair is equivalent to the canonical MO with Mulliken label 1b1, the nO(σ) lone pair is not quite equivalent to the canonical MO of Mulliken label 2a1, since the fully delocalized orbital includes mixing with the in-phase symmetry-adapted linear combination of hydrogen 1s orbitals, making it slightly bonding in character, rather than strictly nonbonding. ↑ Anslyn, E. V.; Dougherty, D. A. (2006). Modern Physical Organic Chemistry. Sausalito, CA: University Science Books. pp. 41. ISBN 978-1-891389-31-3. ↑ Claus, Allen D.; Neisen, Stephen F.; Ayoub, Mohamed; Moore, John W.; Landis, Clark R.; Weinhold, Frank (2014-10-08). "Rabbit-ears hybrids, VSEPR sterics, and other orbital anachronisms". Chemistry Education Research and Practice. 15 (4): 417–434. doi:10.1039/C4RP00057A. ISSN 1756-1108. ↑ Hiberty, Philippe C.; Danovich, David; Shaik, Sason (2015-07-07). "Comment on "Rabbit-ears hybrids, VSEPR sterics, and other orbital anachronisms". A reply to a criticism". Chemistry Education Research and Practice. 16 (3): 689–693. doi:10.1039/C4RP00245H. S2CID 143730926. Retrieved from "

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