I'm not a robot

 FIELDS OF STUDY: Physical Chemistry; Inorganic Chemistry; Organic ChemistryABSTRACTThe visual representation of electron distribution known as Lewis structures and the Lewis dot diagram are discussed. The method represents the valence electrons in atoms and illustrates their combination in shared pairs to form bonds between atoms. Lewis structures are a simple means of relating molecular structure to valence.Gilbert N. Lewis and Valence TheoryIn 1916, American chemist Gilbert N. Lewis (18751946) published "The Atom and the Molecule," his first paper on the role of valence in chemical bonding. Electrons had been identified as charged subatomic particles just nineteen years earlier, and protons were about to be discovered the following year, while the existence of neutrons as the third subatomic particle would not be demonstrated until 1932. The modern theory of atomic structure was therefore in its very early stages and had not yet been defined by quantum mechanics, although the basic underlying mathematical principles had been developed. Any well-defined concept of atomic orbitals and the wave-particle duality of electrons (that is, their tendency to behave as both particles and waves) was still very controversial and very much open to discussion. These are all concepts now taken for granted by chemists and physicists the world over, but in Lewiss time, they were very new ideas. The theoretical aspects did not yet have the support of experimental evidence. The electronic measurement devices to which he had access were crude by modern standards, predating the invention of the transistor by more than thirty years. Accordingly, explanations of chemical behavior based on electron shells and orbitals were subject to intense questioning by traditional chemists.In this environment, Lewis proposed that atoms with an atomic mass greater than that of helium have inner shells of electrons with the same distribution as in the noble gas preceding them in the periodic table (as the six noble gases all have their outer electron shells filled and are therefore chemically inert). A sodium atom, for example, would have all of its electrons but one in the same configuration as the electrons in a neon atom; the extra electron would lie outside of these inner shells of electrons, in the outermost, or valence, shell. The natural corollary of this hypothesis was that the valence electrons of an atom could be easily given up to form a positively charged ion, or cation, with all of its electron shells filled in the same way as those of the preceding noble-gas element. Conversely, electrons could be added from other atoms to form a negatively charged ion, or anion, with all of its electron shells filled in the same way as those of the next noble-gas element in the periodic table.Lewis Dot StructuresAccording to the modern theory of atomic structure, electrons are arranged in shells about the nucleus of an atom. The first shell can hold up to two electrons, the second shell can hold a total of eight, the third shell a total of eighteen, and the fourth shell a total of thirty-two; while subsequent shells could theoretically hold more, in practice, the fifth, sixth, and seventh electron shells also hold a maximum of thirty-two electrons. The outermost shell of an atom, however, only ever has a maximum of eight electrons (at least in theorytransition metals can have more), beyond which the next shell starts filling. For example, a calcium atom has twenty electrons, but it has two electrons in the first shell, eight in the second, eight in the third, and two in the fourth, rather than ten in the third and none in the fourth. The tendency of atoms to combine or form ions in such a way that they have eight electrons in their valence shells is known as the octet rule.In a Lewis dot diagram of an atom, the chemical symbol of the element is shown, representing the nucleus and the inner electron shells, and is surrounded by up to eight dots, generally in pairs, representing the electrons in the atoms outermost shell. For example, using calcium again, the Lewis diagram of this element consists of the element symbol, "Ca," with two dots next to it, representing the two electrons in the fourth shell.In a Lewis structure of a molecule, the chemical symbols of the constituent elements are shown connected by lines in place of dots, each one representing the shared pair of electrons that forms a covalent bond between the atoms; a single line represents a single bond, while double and triple lines represent multiple bonds. (Ionic bonds are represented differently; an ionic compound is depicted as adjacent but separate ions.) Any electrons not involved in chemical bonds, be they unpaired electrons or lone pairs, are still shown as dots next to their respective elements. There are variations on this system, some of which retain the circles representing the atomic orbitals, others of which use the dots in place of lines to show the bonds between atoms in a compound; however, the system described here is the most widely used, as it permits a comprehensible two-dimensional representation of the molecule and the bond system it contains.The Cubical AtomLewiss 1916 paper introduced several important ideas: the sharing of electrons to form a covalent bond, the transfer of electrons from one atom to another to form an ionic bond, the octet rule, and of course his dot diagram. He also proposed that, counter to the planetary model of the atom introduced by Niels Bohr (18851962) in 1913, atoms were in fact cubical in shape, with valence electrons positioned at some or all of the cubes eight corners. Two cubical atoms could form a single covalent bond by sharing a single edge so that they had two corners in common, while a double covalent bond was formed by the atoms sharing a full face, giving them four corners in common. Because this model could not account for triple bondstwo cubes cannot share more than four corners at one timeLewis suggested that in some cases, the electrons of an atom would rearrange themselves from a cubical to a tetrahedral shape (a three-sided pyramid) with two electrons at each corner, allowing two atoms to share six electrons by sharing a single face.While the cubical model of the atom was consistent with Lewiss valence theory, it never found widespread acceptance, although Irving Langmuir (18811957) built on Lewiss ideas to further refine valence theory and propose his own model of atomic structure. Lewiss atomic model, like all other contemporary models, was eventually disproved and replaced by the quantum mechanical model, which was superficially similar to Bohrs model but incorporated the idea of wave-particle duality and defined electron orbitals as the areas of the atom with the greatest probability of containing a given electron, rather than defined paths for electrons to follow around the nucleus.PRINCIPAL TERMScovalent bond: a type of chemical bond in which electrons are shared between two adjacent atoms.electron: a fundamental subatomic particle with a single negative electrical charge, found in a large, diffuse cloud around the nucleus.lone pair: two valence electrons that share an orbital and are not involved in the formation of a chemical bond; also called a nonbonding pair.multiple bond: a bond formed by two atoms sharing two or more electron pairs; includes double bonds and triple bonds.shared pair: the two electrons shared between two atoms in a normal covalent bond.BibliographyAbbott, David, ed. The Biographical Dictionary of Scientists: Chemists. New York: Bedrick, 1983. Print.Askeland, Donald R., and Pradeep P. Phul. The Science and Engineering of Materials. 5th ed. Toronto: Nelson, 2006. Print.Douglas, Bodie Eugene, Darl Hamilton McDaniel, and John J. Alexander. Concepts and Models of Inorganic Chemistry. 3rd ed. New York: Wiley, 1994. Print.Jones, Mark M., et al. Chemistry and Society. 5th ed. Philadelphia: Saunders Coll., 1987. Print.Mackay, K. M., R. A. Mackay, and W. Henderson. Introduction to Modern Inorganic Chemistry. 6th ed. Cheltenham: Nelson, 2002. Print.Myers, Richard. The Basics of Chemistry. Westport: Greenwood, 2003. Print. Lewis structure, also known as Lewis dot structure or electron dot structure, is a simple and straightforward way of representing the outermost electron shell in a chemical species like an atom, ion, or molecule. It shows how electrons are positioned around the atoms either as lone pairs or in a chemical bond, typically a covalent bond or a coordinate covalent bond. Dots represent the lone pairs of electrons, and lines represent the bonded electrons. Since an ion has a charge, the charge is placed as a superscript [1-4]. Lewis Structure Lewis structure is named after American chemist Gilbert Lewis. In 1916, he introduced valence bond theory, dot structures, and dashed lines to represent chemical bonds. The octet rule is used to draw the Lewis structure of a compound. This rule applies to the outmost shell of an atom, which is also known as the valence shell. The electrons in the valence shell are known as valence electrons. The number of valence electrons of an element can be determined from the periodic table. All atoms gain or lose valence electrons to achieve a stable state, i.e., a state with the lowest potential energy. Therefore, atoms tend to have the desired number of electrons in their outermost shell that will enable them to achieve stability. By losing or gaining electrons, atoms take the electronic configuration of their neighboring inert gas atoms. Except for helium, which has 2 electrons, all inert gases have 8 electrons in their outermost shells. This rule that allows atoms to have up to 8 electrons in their valence shell is the octet rule [2-3]. A Lewis structure consists of the following parts [4]. 1. Element symbol 2. Dots representing the lone pairs The element symbol, together with the dots, is called the Lewis symbol. 3. Lines indicating chemical bonds single line for a single or sigma bond, double line for a double bond (one sigma bond and one pi bond), and triple line for a triple bond (one sigma bond and two pi bonds) 4. Brackets around the molecule with a charge placed as a superscript (This point applies only to charged particles) Let us take the example of the covalent compound nitrate ion (NO3) and learn the steps and rules for drawing the Lewis structure of this polyatomic ion [4-9]. Step 1: We start by calculating the total number of valence electrons using the periodic table. First, find this number for each atom of the molecule. Then, add the individual numbers. Finally, to this number, add the electrons coming from the charge of the molecule. The number of valence electrons of the atoms in NO3 is as follows. Nitrogen (N): 5 Oxygen (O): 6 x 3 = 18 Negative charge (-): 1 Total number of valence electrons: 5 + 18 + 1 = 24 Step 2: Next, we draw the skeleton structure of the molecule by connecting the central atom to other atoms by single lines. The central atom is the one with the maximum number of bonding sites. From step 1, we found that nitrogen has 5 electrons and requires 3 more to complete the octet. On the other hand, oxygen has 6 electrons and requires 2 more. Hence, nitrogen with more bonding possibilities is the central atom. Step 3: Next, we will determine the number of electrons needed to complete the octet of each atom. Since each chemical bond comprises 2 electrons, 6 of the 24 electrons will occupy the three bonds in the skeleton. The remaining 18 electrons will go to the terminal atoms, which are the oxygen atoms. These non-bonding electrons are lone pairs. Step 4: Next, we will check if all the octets are filled or not. If they are not, use a lone pair from one of the terminal atoms and form a bonding pair. In this case, one oxygen atom will donate its lone pair to the nitrogen and form a double bond. Step 5: Next, we calculate the formal charge of each atom. The following formula gives the formal charge. qf = V N B/2 where, qf: Formal charge V: Number of valence electrons of the atom N: Number of non-bonding electrons or lone pairs B: Total number of electrons shared in bonds with other atoms Using the formula, the formal charges of nitrogen and oxygen atoms are as follows. Oxygen Atoms with single bonds: 6 6 2/2 = -1 The atom with a double bond: 6 4 4/2 = 0 Nitrogen Only nitrogen atom: 5 0 8/2 = +1 Add all the formal charges to obtain the formal charge of the molecule. -1 + (-1) + 0 + 1 = -1 Hence, the net charge on the nitrate ion is -1. Step 6: In step 5, we have seen the Lewis structure of the nitrate ion. However, this is not the simplest form. To show the simplest Lewis structure, draw a bracket around the molecule and indicate the overall charge as a superscript. The following image summarizes the above steps. How to Draw Lewis Structures Lewis Structure Examples The above image shows the Lewis structures of some common covalent compounds and ions. Learning Objectives To draw Lewis Structures for molecules and polyatomic ionswith one central atom. A Lewis structure is a way to show how atoms share electrons when they form a molecule. Lewis structures show all of the valence electrons in an atom or molecule. The valence electrons are the electrons in the outermost shell.For representative elements, the number of valence electrons equals the group number on the periodic table.To draw the Lewis structure of an atom, write the symbol of the atom and draw dots around it to represent the valence electrons. Note that hydrogen is often shown in both group 1A and group 7A, but it has one valence electron never seven. Also, helium is shown in group 8A, but it only has two valence electrons. Nonmetals can form a chemical bond by sharing two electrons. Each atom contributes one electron to the bond. For example, two hydrogen atoms can form a bond, producing a molecule of H2. Using Lewis structures, we can represent this as follows: Two fluorine atoms can form a molecule of F2 in the same fashion. Note that each atom must contribute one electron to the bond. Atoms can form more than one bond. In a water molecule, an oxygen atom forms two bonds, one to each hydrogen atom. Chemists normally represent a bond using a line instead of two dots. The structures of H2, F2, and H2O would usually be drawn as follows: Only the bonding electrons are shown using lines.Nonbonding electrons are always shown using dots. The bonds on the previous sectionare called single bonds.Each bond contains two electrons (one bonding pair). A pair of atoms can also share four electrons or six electrons. If the atoms share four electrons, the bond is called a double bond.For example, the bond in O2 is a double bond. A double bond is normally depicted with two parallel lines. If the atoms share six electrons, the bond is called a triple bond. Triple bonds are rather rare, but the bond in N2 is a triple bond. A triple bond is depicted with three parallel lines. In Chem 101A, we will focus on drawing Lewis structures of molecules and polyatomic ions that have one central atom with several other atoms attached to it. Such molecules are very common, and they provide a foundation for understanding structures of more complex molecules. To begin, you must know two essential rules for drawing Lewis structures. For example, the basic arrangements of the atoms in SO3, NH4+, and PCl5 are: Do not ever draw a structure like the ones below! None of the molecules above actually look like the structures here.For example, the actual shape of SO3 is: The following procedure will give you the correct Lewis structure for any molecule or polyatomic ion that has one central atom. Step 1: Figure out how many electrons the molecule must have, based on the number of valence electrons in each atom.When drawing the structure of an ion, be sure to add/subtract electrons to account for the charge. Step 2: Connect the atoms to each other with single bonds to form a skeleton structure.Be sure that you follow rule 1 in the previous section. Step 3: Add enough electrons (dots) to the outer atoms to give each of them a total of eight electrons around them.(Exception: do not add electrons to hydrogen atoms.) This tendency of atoms to have eight electrons around them is called the octet rule. Step 4: Count the electrons in your structure.If you need to add any more based on your count in step 1, add them to the central atom.The electron count in your final answer must match the count from step 1. Step 5: If the central atom has 8 or more electrons around it, youre finished. If the central atom has fewer than 8 electrons around it, but all of the surrounding atoms are from group 7A, youre finished. Otherwise, move a nonbonding electron pair from an outer atom to a bond (i.e. make a double bond). If the central atom now has 8 electrons around it, youre finished. Otherwise, repeat this process until the central atom has 8 electrons. Step 1) Figure out how many electrons the molecule must have. Carbon has 4 valence electrons Each oxygen has 6 valence electrons The -2 charge means that there are 2 extra electrons Total: 4 + (3 6) + 2 = 24 electrons The final answer MUST have this number of electrons! Step 2) Attach the atoms to each other using single bonds (draw the skeleton structure) Step 3) Add electrons to all outer atoms (except H) to complete their octets. The outer atoms are the oxygen atoms here. Each outer atom needs three electron pairs, since it already has one bonding pair. (1 line = 2 electrons) Step 4) Count the electrons in the structure. This structure has 24 electrons. 3 lines = 6 bonding electrons 18 dots = 18 nonbonding electrons total = 24 electrons Does this match the count you got in step 1? CO32: should have 24 electrons (from step 1) Our structure has 24 electrons check! It MATCHES! Step 5) If the central atom has 8 or more electrons, youre done. The carbon atom has only 6 electrons around it, so we arent finished yet If the central atom has fewer than 8 electrons, but all of the outer atoms are in group 7A, youre done. The outer atoms are oxygen atoms, and oxygen is in group 6A, so we arent finished yet Otherwise, create an extra bond by changing one of the nonbonding pairs into a bonding pair. (Exception: dont do this if all of the outer atoms are from group 7A.) Notice that this step does not change the total number of electrons in our structure. It just moves 2 electrons to a different location. If the central atom still has fewer than 8 electrons around it, do this as many more times as you need.But do not go beyond 8 electrons on the central atom.When you reach 8 electrons, youre done. Chemists often draw square brackets around the structure of a polyatomic ion and write the charge outside the brackets, like this: Note: we could have put the double bond in two other locations.Any of the three options is fine; you only need to draw one of them. (If youre wondering what about resonance?? well get to that later on.) Next, well look at three molecules side-by-side.The molecules are BF3, PF3, and BrF3, all of which have a central atom bonded to three fluorine atoms. Step 1) Figure out how many electrons each molecule must have. BF3: 3 + 7 + 7 + 7 = 24 electrons PF3: 5 + 7 + 7 + 7 = 26 electrons BrF3: 7 + 7 + 7 + 7 = 28 electrons The final answers MUST have these numbers of electrons! Step 2) Attach the atoms to each other using single bonds (draw the skeleton structure). Step 3) Add electrons to the outer atoms, to complete their octets. Each outer atom needs three electron pairs. Step 4) Count the electrons in each structure. Each of these structures has 24 electrons. 3 lines = 6 bonding electrons 18 dots = 18 nonbonding electron total = 24 electrons Do these match the counts you got in step 1? BF3: should have 24 electrons (from step 1) Our structure has 24 electrons MATCHES PF3: should have 26 electrons (from step 1) Our structure has 24 electrons WE NEED TO ADD 2 MORE ELECTRONS BrF3: should have 28 electrons (from step 1) Our structure has 24 electrons WE NEED TO ADD 4 MORE ELECTRONS If the counts do not match, add the remaining electrons to the central atom. Step 5) If the central atom has 8 or more electrons around it, youre done. If the central atom has fewer than 8 electrons around it, but all of the outer atoms are in group 7A, youre done Note: these six elements from group 7Aare called halogens:F, Cl, Br, I At, Tn. Youll only see the first four of them in chemical compounds; the last two are extremely radioactive. The central atom in a structure often violates the octet rule.Electron-deficient atoms are rare, but expanded octets are fairly common with elements in the 3rd row and beyond. In BF3, the central atom only has 6 electrons around it. We say that the boron atom is electron deficient. In BrF3, the central atom has 10 electrons around it. We say that the bromine atom has an expanded octet. Note that it is also quite common for the central atom to makemore than four bonds.Here are a couple of examples: Dont panic if you see such molecules. Just follow the rules for drawing Lewis structures, and youll get them right!However, you must note that outer atoms never violate the octet rule. If we draw the Lewis structure for PO43 ion using the rules youve seen earlier, we come up with the structure below.It has the correct number of electrons (32), and every atom satisfies the octet rule. However, many textbooks (and websites) insist that the structure below is a better one, even though the phosphorus atom has ten electrons around it: The first structure follows the rules for drawing structures. So why do chemists add a double bond? To understand, we need the formal charges on the atoms in each structure Many chemists prefer the second structure because it gives us a couple of zeroes.However, the phosphorus atom in the second structure violates the octet rule. So which structure is best?? The answer is that either structure is legitimate.Both structures give us all of the information we need about phosphate ion; they allow us to predict the shape of the molecule, the angles between the bonds, and whether the molecule is polar. So which structure should YOU draw on a test? Be sure to check with your instructor, but most willaccept either one.However, the first structure is easier to figure out, because its the structure you produce when you follow the provided rules.The second structure requires more work.Therefore, we recommend that when you draw a structure that satisfies the octet rule, you stop there without adding more bonds. WARNING: Some students come to Chem 101A having been taught to draw Lewis structures with extra double bonds. If youre familiar with Lewis structures and you like the extra bonds congratulations! You have less to learn you already know all of this stuff. BUT If you draw structures with extra bonds, they have to be correct.For example, a structure for phosphate that has two double bonds is not acceptable. Also, a structure for nitrate ion (NO3) that has two double bonds is not acceptable, even though it gives you more zeroes.The structure with one double bond is the only acceptable one. If you dont know why the structures Ive labeled unacceptable are not allowed, dont risk losing points by adding extra bonds when the central atom already has eight electrons. Generally, you are told which atom is the central bonding atom in the molecule. However, if it is unclear, or if you are asked to decide which atom is central, formal charges can be used to decide. The following is an example. Example: calculating the formal charges in HCN and HNC For the arrangement HCN,the Lewis structure: HC\(\equiv\)N: The formal charges work out as follows: For the arrangement HNC,the Lewis structure: HN\(\equiv\)C: The formal charges work out as follows: Both Lewis structures have a net formal charge of zero, but note thatthe formal charges on the first structure are all zero! Thus the first Lewis structureis predicted to be more stable, and it is, in fact, the structure observed experimentally. In general, the closer the formal charges are to zero, the more stable the structure. Remember, though, that formal charges donotrepresent the actual charges on atoms in a molecule or ion. They are used simply as a book-keeping methodfor predicting the most stable Lewis structure for a compound. James Armstrong, City College of San Francisco Torrey Glenn, City College of San Francisco Here are the steps to draw a Lewis structure. The example is for the nitrate ion.A Lewis structure is a diagram that shows the chemical bonds between atoms in a molecule and the valence electrons or lone pairs of electrons. The diagram is also called a Lewis dot diagram, Lewis dot formula, or electron dot diagram. Lewis structures take their name from Gilbert N. Lewis, who introduced valence bond theory and dot structures in the 1916 article The Atom and the Molecule.A Lewis structure shows how electrons are arranged around atoms, but it doesnt explain how the electrons are shared between atoms, how chemical bonds form, or what the geometry of a molecule is. Here is how to draw a Lewis structure, with examples and a look at both the importance and limitations of the diagrams.Lewis structures are drawn for molecules and complexes. A Lewis structure consists of the following parts:Element symbolsDots that indicate valence electronsLines that indicate chemical bonds (one line for a single bond, two for a double bond, etc.)The dots and lines satisfy the octet rule.If the structure carries a net charge, brackets enclose it and the charge is listed in the upper righthand cornerNote: Sometimes the terms Lewis structure and electron dot structure are used interchangeably. Technically, they are a bit different. A Lewis structure uses lines to indicate chemical bonds, while an electron dot structure only uses dots.There are only a few steps to draw a Lewis structure, but it can take some trial and error to get it right.Find the total number of valence electrons for all atoms in the molecule. For a neutral molecule, this is the sum of the valence electrons in each atom. The number of valence electrons for an element is usually the same as its group number of on the periodic table (except for helium and the metals). If the molecule has a charge, subtract one electron for each positive charge or add one electron for each negative charge. For example, for NO3, you have 5 electrons for the nitrogen atom and 3 x 6 = 18 electrons for the oxygen atoms, plus one valence electron for the net charge, giving a total of 24 valence electrons (5 + 18 + 1).Draw the skeleton structure of the molecule. At this point, assume the atoms are connected by single bonds. Usually, the atom that has the most bonding sites is the central atom (so carbon would be central over oxygen).Determine how many electrons are needed to satisfy the octet rule. The valence electron shell of hydrogen and helium fill with 2 electrons. For other atoms, up to period 4 of the periodic table, the valence shell fills with 8 electrons. Each chemical bond requires two electrons, so use two valence electrons to form each bond between atoms in the skeleton structure. For NO3, 6 electrons were used to draw the single bonds for the skeleton. So, 18 electrons remain. Starting with the most electronegative atom, distribute these electrons to try to fill the octets of the atoms.Distribute the remaining valence electrons. Draw these non-bonding electrons as dots around the atoms to satisfy the octet rule.Draw the chemical bonds in the molecule. If all of the octets arent filled, make double bonds or triple bonds. To do this, use a lone pair of electrons on an electronegative atom and make it into a bonding pair shared with an electropositive atom that lacks electrons.Check to make sure you have the lowest formal charge for each atom. Dont violate the octet rule. The formal charge is the number of valence electrons, minus half the number of bonding electrons, minus the number of lone electrons. So, for each single-bonded oxygen its 6 1 6 = -1; for nitrogen its 5 4 0 = +1; for the double-bonded oxygen its 6 2 4 = 0. There are two single-bonded oxygen atoms, one nitrogen, and one double-bonded oxygen, so the net formal charge is -1 + -1 + 1 + 0 = -1. Either indicate the formal charges separately or else draw a bracket around the structure and add or -1 as a superscript.A Lewis structure includes lines for covalent chemical bonds and dots for valence electrons or lone electron pairs.There is more than one right way to draw a Lewis structure. If you are drawing the structures for a chemistry class, be sure to know what your instructor expects. For example, some chemists prefer to see skeletal structures that do no show any geometry, while other prefer to see shapes (e.g., the bent shape of water, with nonbonding electron pairs at an angle on one side of the oxygen atom). Some like to see atoms and their electrons in color (e.g., oxygen and its electrons in red, carbon and its atoms in black).Lewis structures help describe valence, chemical bonding, and oxidation states because many atoms fill or half-fill their valence shell. The behavior described by the structures closely approximates real behavior of lighter elements, which have eight valence electrons. So, they are particularly helpful in organic chemistry and biochemistry, which relies on the behavior of carbon, hydrogen, and oxygen. Although Lewis structures do not necessarily show geometry, they are used to predict geometry, reactivity, and polarity.While useful for some applications, Lewis structures arent perfect. They dont work well when molecules contain atoms with more than eight valence electrons, such as the lanthanides and actinides. Inorganic and organometallic compounds employ bonding schemes beyond those described by Lewis structures. In particular, molecular orbitals may be fully delocalized. Lewis structures do not account for aromaticity. Even with lighter molecules (O2, ClO2, NO), the predicted structures differ from real behavior enough that Lewis structures might lead to incorrect predictions about bond length, magnetic properties, and bond orders.IUPAC (1997). Lewis formula. Compendium of Chemical Terminology (the Gold Book) (2nd ed.). Blackwell Scientific Publications. ISBN 0-9678550-9-8.Lewis, G. N. (1916), The Atom and the Molecule. J. Am. Chem. Soc. 38 (4): 76285. doi:10.1021/ja02261a002Miburo, Barnabe B. (1993). Simplified Lewis Structure Drawing for Non-science Majors. J. Chem. Educ. 75 (3): 317. doi:10.1021/ed075p317Zumdahl, S. (2005) Chemical Principles. Houghton-Mifflin. ISBN 0-618-37206-7.Related Posts Welcome to the Science Notes and Projects site, where youll find experiments, projects, and notes for all scientific disciplines and educational levels. We write science articles and make all kinds of printable periodic tables and science graphics.Heres the Science Notes periodic table with the official names and data for all 118 elements.Heres a comprehensive chemistry dictionary. There is a separate page for each letter so that you dont have to scroll forever. Many terms link to articles on the topic.Here are two lists for you. One covers things that glow under black light (fluorescent or phosphorescent), while the other covers things that glow when the lights go out and its truly dark. Some items produce luminescence both ways:Things That Glow in the DarkThings That Glow Under Black LightSometimes its just easier to have a printout of important information. These tables are available as PDF files. You can save these and print them out at your convenience:Here are some fun science projects to try:Add some humor to your life with a collection of science and math jokes. Spend a little time laughing (or groaning) at the lighter side of science. Post a reply to tell your own science joke!Chemistry JokesMath JokesPhysics JokesEngineer JokesLearn to love the triangle. Trigonometry is one of the scientists basic tools in their mathematics toolbox. Here are a couple of useful printable reference sheets with trig relationships, special triangles and trig identities. Got requests for trig or stats content? Let us know. We both have degrees in math.Right TrianglesTrigonometry FunctionsSpecial TrianglesTrig Tables SIN, COS and TAN values PDFTrig Identities Study SheetMetals, Metalloids and NonmetalsThis periodic table shows the elements and locations of the metals, metalloids and nonmetals. It also lists the common properties of metals, metalloids and nonmetals. You can print out the table the same as any of the periodic tables on the site.List of MetalsThis is a list of the elements considered to be the metal elements along with a periodic table highlighting these elements and their common properties.List of MetalloidsThis is a list of the elements considered to be the metalloid or semimetal elements. It also contains a periodic table highlighting the metalloids and their common properties.List of NonmetalsThis is a list of all the nonmetal elements, their location on the periodic table and the properties of a nonmetal.Here are some new science projects, plus some old favorites:Weve got a bunchcrystal-growing projects for you. Instructions range from crystals grown with home chemicals to easy lab chemical crystals to try. Written By Sushmita Rout Last Modified 25-01-2023 Lewis Structures:As valence electrons are significant to an atoms reactivity, it is essential to represent it by simple diagrams.Lewis Structuresare pictorial representations of molecules in which the valence electrons present in an atom are represented as dots. Hence, these structures are also known as electron dot diagrams. The Lewis structure was named after Gilbert N. Lewis, who introduced it in his 1916 article called The Atom and the Molecule. Lewis dot structures are extensions of the concept of the electron dot diagram. This structure can be drawn for any covalently bonded molecule and coordination compound. Read on to learn more exciting facts about Lewis structures and how to draw SO2 Lewis Structure, NH3 Lewis Structure, CO2 Lewis Structure and more. Atoms are made up of neutrons, protons \((+)\) and electrons \((-)\). The nucleus of an atom is made up of positive protons and neutral neutrons. Electrons are found moving around the nucleus in energy shells. The shells can be numbered numerically as \(1,2,3\) so on or by using letters like \({\rm{K}},{\rm{L}},{\rm{M}}\) so on. This means the energy shell K is denoted as \(1\) and has the lowest energy. The orbits are arranged in increasing energy levels, with the lowest energy shell being closer to the nucleus and the highest energy shell being farthest from the nucleus. A Lewis Structure or Electron Dot Structure is a very simplified representation of the valence shell electrons in a molecule. It denotes the way the valence electrons are arranged around the individual atoms in a molecule. The Lewis structure was named afterGilbert N. Lewis, who introduced it in his \(1916\) article The Atom and the Molecule. In a Lewis Structure, electrons are represented as dots surrounding the central metal atom. The central metal is denoted by using its chemical symbol from the Periodic Table. In Lewis Structures, a line is used to represent the bonding electrons between two combining atoms. The Lewis Structure also denotes the number of lone pairs of electrons present around the central atom. Lewis Structures can be drawn for ionic, covalent and coordination compounds. However, we can draw it only if we know the molecular formula of the compound. It is also helpful to represent the atomic state of elements. We represent an electron dot structure by using the symbol of an element. The dots that represent the valence electrons are added to the chemical symbol of an element in a clockwise manner. We cannot put the dots anywhere around the symbol. Instead, we need to imagine a square around the chemical symbol. The dots should be neatly drawn on the four sides of the square with no more than two electrons on each side. The maximum number of dots can be eight. This is because the maximum number of valence electrons can be only eight, thereby satisfying the octet rule.The octet rule is significant to a Lewis dot structure. The octet rule states that atoms tend to gain, lose, or share valence electrons to attain noble gas configuration, which imparts stability to the atom. Only the s and p electrons are involved in the octet rule; the d and f electrons are not considered. The steps needed to draw them are: The total number of valence electrons present in the molecule of the compound is calculated by adding the individual valence electrons of each atom. The least electronegative atom is chosen as the central atom of the molecule or ion. The central metal atom is the one to which all the other atoms will be bonded. The central atom to be chosen should possess the least subscript in a given molecule. A single bond represents \(2\) valence electrons, one from each atom contributing to bond formation. A line represents a single bond. The number of valence electrons used for bonding in step \(3\) is subtracted from the total number of valence electrons calculated in step \(1\). The remaining electrons are assigned to each atom as lone pair of electrons. Turn all lone pairs into double or triple bonds to satisfy the octet configuration of each combining atom. For an anion, we need to add extraelectronsto the dot structure. The number of an extra electron that is to be added is always equal to the charge on the anion.For a cation, we need to subtract electrons from the total count. The number of extra electrons that is to be subtracted is always equal to the charge on the cation. Let us take the example of Sulphur trioxide \({\rm{S}}{{\rm{O}}_{\rm{3}}}\) Step 1 Calculation of the total number of valence electrons in sulfur trioxide molecule. Both oxygen and Sulphur belong to group \(16\) of the periodic table. Hence, Sulphur and oxygen contain \(6\) valence electrons each. So in \({\rm{S}}{{\rm{O}}_{\rm{3}}}\) we have \({\rm{S = 6 \times 1 = 6}}\) valence electrons \({\rm{O = 6 \times 3 = 18}}\) valence electrons Hence, there are a total of \(24\) valence electrons in \({\rm{S}}{{\rm{O}}_{\rm{3}}}\) molecule. Step 2 Determination of the central metal atom. In \({\rm{S}}{{\rm{O}}_{\rm{3}}}\), Sulphur being less electronegative than oxygen, it is chosen as the central metal atom. Also, Sulphur has the lowest subscript. We can represent sulfur trioxide as below- Step 3- Draw single bonds to the central atom. Step 4 Calculation of lone pairs of electrons. In total, we had \(24\) valence electrons as calculated in Step 1. There are three single bonds, so we have used \(3\) pairs or \(6\) valence electrons out of \(24\). The remaining valence electrons are\(18\) in the count. These remaining \(18\) valence electrons are used as lone pairs on the atom. This is as shown below- Step 5 Satisfying the octet configuration for each of the participating atoms. In \({\rm{S}}{{\rm{O}}_{\rm{3}}}\), the Sulphur atom does not have an octet configuration. It has only \(6\) valence electrons. So, to complete the octet configuration of Sulphur, one of the three oxygen atoms shares one lone pair of electrons with the Sulphur atom. This results in the formation of a double bond. The steps to draw the Lewis structures of various types of compounds are given below: Oxygen belongs to group \(16\) of the Periodic Table. Hence, oxygen has \(6\) valence electrons. Step 1 Calculating the total valence electrons. In \({{\rm{O}}_2}\), total number of valence electrons \(= 6 2 = 12\) valence electrons Step 2 Determination of the central metal atom. In \({{\rm{O}}_2}\), there are only two oxygen atoms. Hence, we can choose any one of the oxygen atoms as the central metal atom. Step 3 Connecting the participating atoms through single bonds. In \({{\rm{O}}_2}\), \(2\) of the oxygen atoms are connected through single bonds. The structure of the \({{\rm{O}}_2}\) molecule in Step \(3\) is as shown below Step 4 Calculation of lone pair of electrons. In \({{\rm{O}}_2}\), Total number of valence electrons \(= 6 2 = 12\) valence electrons Valence electrons used in making \(1\) single bond \(= 1 2 = 2\) valence electrons Hence, total number of valence electrons left after making a single bond \(= 12 2 = 10 \)valence electrons. These \(10\) valence electrons are distributed over the \(2\) oxygen atoms. This is as shown below- Step 5 Satisfying the octet configuration for each of the participating atoms. In \({{\rm{O}}_2}\), one of the oxygen atoms does not possess the octet configuration. Hence, the octet configuration for this oxygen atom is fulfilled by the other oxygen atom. The other oxygen atom shares its lone pair of electrons with the oxygen atom lacking an octet configuration. This results in the formation of a double bond. The Lewis structure for oxygen molecule is as shown below- In \(\left({{\rm{C}}{{\rm{O}}_2}} \right),\) oxygen belongs to group \(16\) of the Periodic Table, and carbon belongs to group \(14\) of the Periodic Table. Hence, oxygen has \(6\) valence electrons, and carbon has \(4\) valence electrons. Step 1- In \(\left({{\rm{C}}{{\rm{O}}_2}} \right)\) we, have \({\rm{C = 4 \times 1 = 4}}\) valence electrons \(O = 6 \times 2 = 12\) valence electrons Hence, there are \(16\) valence electrons in \(\left({{\rm{C}}{{\rm{O}}_2}} \right)\) molecule. Step 2 Determination of the central metal atom. In \(\left({{\rm{C}}{{\rm{O}}_2}} \right)\), carbon has the lowest subscript. So, carbon is chosen as the central atom. We can represent carbon dioxide as below- Step 3 Connecting the two oxygen atoms through single bonds to the carbon atom. Step 4 Calculation of lone pairs of electrons. Total number of valence electrons \(= 16\) Valence electrons used in bonding \(= 1 \times 2 = 2(1\) single bond \(=2\) electrons). Valence electrons remaining \(= 16 4 = 12\) Hence, \(12\) valence electrons are remaining, distributed as lone pairs over the \(2\) oxygen atoms. Step 5 Satisfying the octet configuration. After step \(4\), it was found that the octet configuration of both the oxygen atoms is satisfied, but the octet configuration of the carbon atom is not satisfied. The oxygen atoms possessing lone pair of electrons share it with the neighbouring carbon atomsThis leads to a double bond between the \(2\) oxygen atoms and carbon atoms pairs, as shown below. In \({\rm{CO}}\), oxygen belongs to group \(16\) of the Periodic Table, and carbon belongs to group \(14\) of the Periodic Table. Hence, oxygen has \(6\) valence electrons, and carbon has \(4\) valence electrons. Step 1- Calculating the total number of valence electrons. In \({\rm{CO}}\) we have- \({\rm{C = 4 \times 1 = 4}}\) valence electrons \({\rm{O = 6 \times 1 = 6}}\) valence electrons Hence, there are \(10\) valence electrons in \({\rm{CO}}\) molecule. Step 2 Determination of the central metal atom. In \({\rm{CO}}\), both carbon and oxygen have the lowest subscript. The carbon atom is less electronegative than the oxygen atom. Hence, carbon is chosen as the central atom. We can represent carbon monoxide as below- Step 3 Connecting the participating atoms (C and O) through a single bond. Step 4 Calculation of lone pairs of electrons. Total number of valence electrons \(= 10\) Valence electrons used in bonding \(= 1 \times 2 = 2(1\) single bond \(=2\) electrons) Valence electrons remaining \(= 10 2 = 8\) Hence, \(8\) valence electrons are remaining, distributed as lone pairs over the oxygen atom and carbon atom. Step 5 Satisfying the octet configuration. After step \(4\), it was found that the octet configuration of the oxygen atom is satisfied but for the carbon atom, it is not satisfied. The oxygen atom possessing \(3\) lone pairs of electrons share two lone pairs with the neighbouring carbon atom. This leads to a triple bond between the carbon atom and oxygen atom, which is shown below. In \(\left({{\rm{N}}{{\rm{H}}_{\rm{3}}}} \right)\), nitrogen belongs to group \(15\) of the Periodic Table, and Hydrogen belongs to group \(1\) of the Periodic Table. Hence, nitrogen has \(5\) valence electrons, and hydrogen has \(1\) valence electron. Step 1- Calculating the total number of valence electrons. In \(\left({{\rm{N}}{{\rm{H}}_{\rm{3}}}} \right)\), we have \({\rm{N = 5 \times 1 = 5}}\) valence electrons \({\rm{H = 1 \times 3 = 3}}\) valence electrons Hence, there are \(8\) valence electrons in \(\left({{\rm{N}}{{\rm{H}}_{\rm{3}}}} \right)\) Molecule. Step 2 Determination of the central metal atom. In \(\left({{\rm{N}}{{\rm{H}}_{\rm{3}}}} \right)\), nitrogen has the lowest subscript. Hence, it is chosen as the central metal atom. We can represent ammonia as below- Step 3 Connecting the participating atoms (N and H) through a single bond. Step 4 Calculation of lone pairs of electrons. Total number of valence electrons \(= 8\) Valence electrons used in bonding \(= 3 2 = 6\) (\(3\) single bond \(= 6\) electrons) Valence electrons remaining \(= 8 6 = 2\) Hence, \(2\) valence electrons are remaining, distributed as lone pair over the nitrogen atom. Step 5 Satisfying the octet configuration. After step \(4\), it was found that the octet configuration of the nitrogen atom and hydrogen atom is satisfied. As the octet configuration of all the participating atoms that is nitrogen and hydrogen is satisfied, there is no need for extra bond formation. The Lewis structure of ammonia is as shown below- The formal charge on an atom in a molecule reflects the electron count associated with the atom compared to the isolated neutral atom. If the atom has given away electrons, it will be \(+ve\), and if it has gained electrons, it will be \(-ve\). No Lewis structure is complete without the formal charge. The goal is to obtain the best electron configuration,i.e.the octet rule and formal charges need to be satisfied.It defines the nature of the bond and position of atoms of the molecule which are connected in the molecule.A Lewis dot structure illustrates the sharing of electrons between atoms in covalent or polar covalent bonds To conclude, we can say that Lewis Structures are of great importance in explaining the arrangement of outer shell valence electrons of an atom or a molecule. Lewis structures help us to visualize the number of valence electrons present around an atom or a molecule. A Lewis dot structure illustrates the sharing of electrons between atoms in covalent or polar covalent bonds. It defines the nature of the bond and position of atoms of the molecule which are connected in the molecule. The central atom is usually the atom with the lowest subscript in the molecular formula and the atom that can form the most bonds. It is also the atom having low electronegativity. If all of the atoms usually form the same number of bonds, the least electronegative atom is usually the central atom. In addition, it gives us an idea about the bond type and the lone pair of electrons present over the participating atoms. In this article, we studied what are the weird lines and dots surrounding an atom and how they are useful in chemistry. The most commonly asked questions about Lewis structure are answered here: Q.1. How do you know where to put the dots on a Lewis structure?Ans: The dots representing the valence electrons are added to the chemical symbol of an element in a clockwise manner. Therefore, we cannot put the dots anywhere around the symbol. Instead, we need to imagine a square around the chemical symbol. The dots should be neatly drawn on the four sides of the square with no more than two electrons on each side. Q.2. What do Lewis structures show?Ans: Lewis structureshelp us to visualize the number of valence electrons present around an atom or a molecule. In addition, it gives us an idea about the bond type and the lone pair of electrons present over the participating atoms. Q.3. What are the limitations of Lewis Structures?Ans: The limitations of Lewis structure are as follows-1. Lewis structures do not account for the aromaticity of the compound.2. Lewis structures do not hold good for the resonance structures.3. Lewis structures do not explain the magnetic behaviour of compounds.4. These structures do not account for the shape and geometry of the molecule. Q.4. What is the central atom in a Lewis structure?Ans: Thecentral atoms usually theatomwith the lowest subscript in the molecular formula and theatomthat can form the most bonds. It is also the atom having low electronegativity. If all of theatomsusually form the same number of bonds, the least electronegativeatomis usually thecentral atom. Q.5. What is the significance of Lewis Structures?Ans: A Lewis dot structure illustrates the sharing of electrons between atoms in covalent or polar covalent bonds. It defines the nature of the bond and position of atoms of the molecule which are connected in the molecule. Learn Formal Charge In Lewis Structure We hope this article is helpful to you. If you have any queries on Lewis Structure, ping us through the comment box below and we will get back to you as soon as possible.

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