
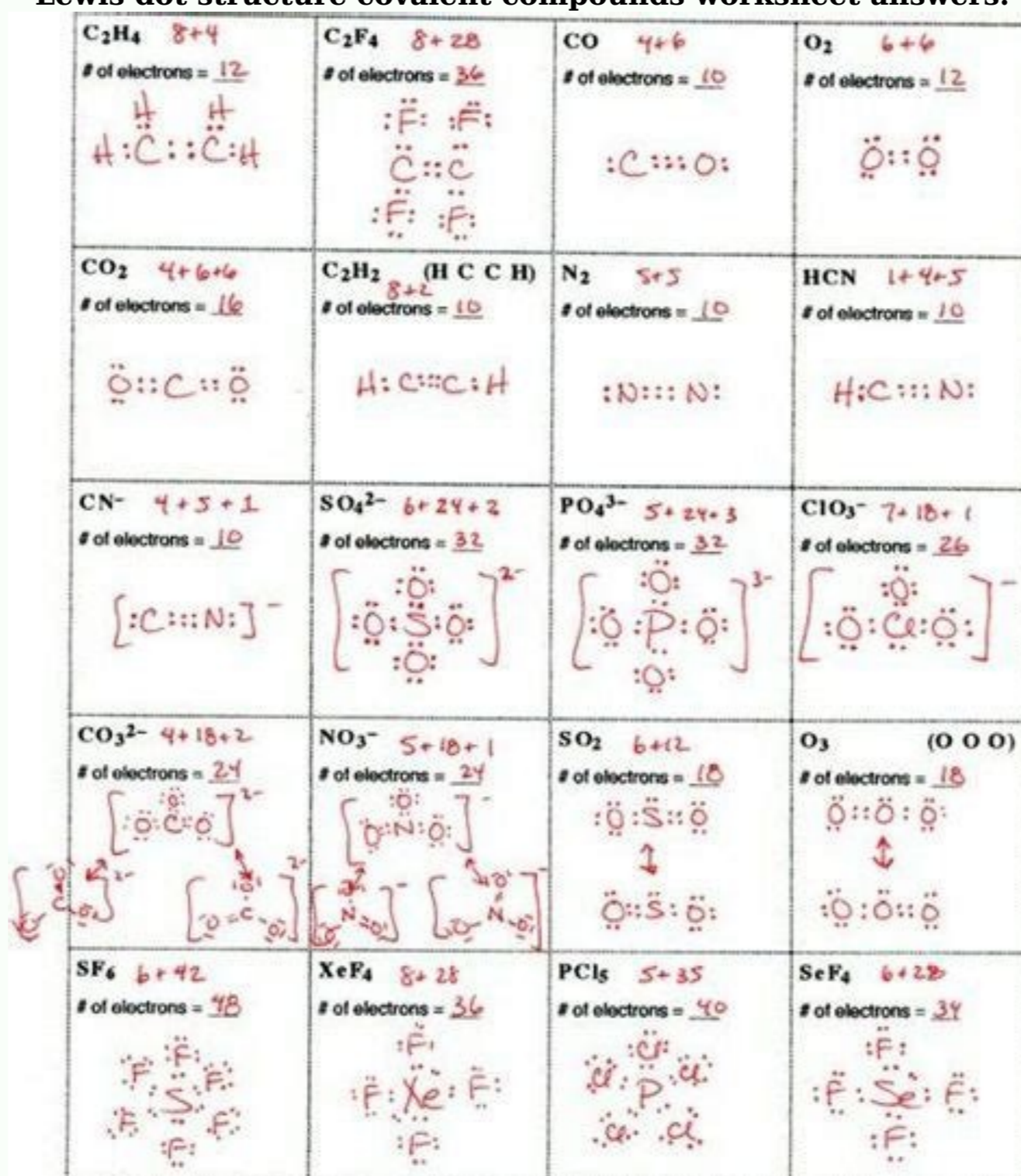


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Ionic and covalent lewis dot structures worksheet

Lewis dot structure for simple ionic and covalent compounds. Lewis dot structure covalent compounds worksheet answers.



Ionic and covalent bonds lewis structure. Use lewis dot structures to show the ionic bonding.

Chapter 1 • Chapter 2 • Chapter 3 • Chapter 4 • Chapter 5 • Chapter 6 • Chapter 7 • Chapter 8 Chapter 9 • Chapter 10 • Chapter 11 • Chapter 12 • Chapter 13 • Chapter 14 Learning Objectives State the octet rule.

Covalent Bonding - Questions	
CH_4 1 Carbon + 4 Hydrogen	
NO_3 1 Nitrogen + 3 Oxygen	
HF 1 Hydrogen + 1 Fluorine	

Define ionic bond. Draw Lewis structures for ionic compounds. In Section 4.7, we demonstrated that ions are formed by losing electrons to make cations, or by gaining electrons to form anions. The astute reader may have noticed something; many of the ions that form have eight electrons in their valence shell. Either atoms gain enough electrons to have eight electrons in the valence shell and become the appropriately charged anion, or they lose the electrons in their original valence shell; the lower shell, now the valence shell, has eight electrons in it, so the atom becomes positively charged.

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For whatever reason, having eight electrons in a valence shell is a particularly energetically stable arrangement of electrons. The octet rule explains the favorable trend of atoms having eight electrons in their valence shell. When atoms form compounds, the octet rule is not always satisfied for all atoms at all times, but it is a very good rule of thumb for understanding the kinds of bonding arrangements that atoms can make. It is not impossible to violate the octet rule. Consider sodium: in its elemental form, it has one valence electron and is stable. It is rather reactive, however, and does not require a lot of energy to remove that electron to make the Na^+ ion. We could remove another electron by adding even more energy to the ion, to make the Na^{2+} ion. However, that requires much more energy than is normally available in chemical reactions, so sodium stops at a $1+$ charge after losing a single electron. It turns out that the Na^+ ion has a complete octet in its new valence shell, the $n = 2$ shell, which satisfies the octet rule. The octet rule is a result of trends in energies and is useful in explaining why atoms form the ions that they do. Now consider an Na atom in the presence of a Cl atom. The two atoms have these Lewis electron dot diagrams and electron configurations: $\text{Na} \cdot$ and $\text{Cl} \cdot$. An electron transfers from the Na atom to the Cl atom: $\text{Na} \cdot + \cdot \text{Cl} \rightarrow \text{Na}^+ + \text{Cl}^-$. Both species now have complete octets, and the electron shells are energetically stable. From basic physics, we know that opposite charges attract. This is what happens to the Na^+ and Cl^- ions: $\text{Na}^+ + \text{Cl}^- \rightarrow \text{NaCl}$. We have written the final formula (the formula for sodium chloride) as per the convention for ionic compounds, without listing the charges explicitly. The attraction between oppositely charged ions is called an ionic bond, and it is one of the main types of chemical bonds in chemistry. Ionic bonds are caused by electrons transferring from one atom to another. In electron transfer, the number of electrons lost must equal the number of electrons gained. We saw this in the formation of NaCl. A similar process occurs between Mg atoms and O atoms, except in this case two electrons are transferred: $\text{Mg} \cdot + \cdot \text{O} \cdot \rightarrow \text{Mg}^{2+} + \text{O}^{2-}$. Remember, in the final formula for the ionic compound, we do not write the charges on the ions. What about when an Na atom interacts with an O atom? The O atom needs two electrons to complete its valence octet, but the Na atom supplies only one electron: $\text{Na} \cdot + \cdot \text{O} \cdot \rightarrow \text{Na}^+ + \text{O}^-$. The O atom still does not have an octet of electrons. What we need is a second Na atom to donate a second electron to the O atom: $2\text{Na} \cdot + \cdot \text{O} \cdot \rightarrow \text{Na}^+ + \text{Na}^+ + \text{O}^{2-}$. These three ions attract each other to give an overall neutral-charged ionic compound, which we write as Na_2O . The need for the number of electrons lost being equal to the number of electrons gained explains why ionic compounds have the ratio of cations to anions that they do. This is required by the law of conservation of matter as well. With arrows, illustrate the transfer of electrons to form calcium chloride from (Ca) atoms and (Cl) atoms. A (Ca) atom has two valence electrons, while a (Cl) atom has seven electrons. A (Cl) atom needs only one more to complete its octet, while (Ca) atoms have two electrons to lose. Thus we need two (Cl) atoms to accept the two electrons from one (Ca) atom. The transfer process looks as follows: The oppositely charged ions attract each other to make $CaCl_2$. With arrows, illustrate the transfer of electrons to form potassium sulfide from (K) atoms and (S) atoms. Answer: The tendency to form species that have eight electrons in the valence shell is called the octet rule. The attraction of oppositely charged ions caused by electron transfer is called an ionic bond. The strength of ionic bonding depends on the magnitude of the charges and the sizes of the ions. The following procedure can be used to construct Lewis electron structures for more complex molecules and ions. Add together the valence electrons from each atom. (Recall that the number of valence electrons is indicated by the position of the element in the periodic table.) If the species is a polyatomic ion, remember to add or subtract the number of electrons necessary to give the total charge on the ion. For CO_3^{2-} , for example, we add two electrons to the total because of the -2 charge. 2. Arrange the atoms to show specific connections. When there is a central atom, it is usually the least electronegative element in the compound. Chemists usually list this central atom first in the chemical formula (as in CCl_4 and CO_3^{2-} , which both have C as the central atom), which is another clue to the compound's structure. Hydrogen and the halogens are almost always connected to only one other atom, so they are usually terminal rather than central. In H_2O , for example, there is a bonding pair of electrons between oxygen and each hydrogen. 4.

Subshell	Electron Structure	Electron Structure	Subshell	Electron Structure	Electron Structure
1s			2s		
2s			2p		
3s			3p		
4s			3d		

Beginning with the terminal atoms, add enough electrons to each atom to give each atom an octet (two for hydrogen). These electrons will usually be lone pairs.

Name _____ Date _____

Conceptual Chemistry

How Atoms Bond
Electron-Dot Structures

Electron-dot structures tell us how atoms tend to bond with other atoms. Carbon's electron-dot structure, for example, shows 4 unpaired valence electrons. Each hydrogen atom has 1 unpaired electron. Unpaired electrons from different atoms can pair up resulting in a bond. For carbon and hydrogen, we have the following, which creates the molecule methane, CH₄.

Use this idea to draw the electron-dot structures of the products indicated by name:

Ammonia, NH₃

Chlorine, Cl₂

Hydrazine, N₂H₄

Nitrogen, N₂ + Hydrogen, H₂

Hydrogen peroxide, H₂O₂

Oxygen, O₂ + Hydrogen, H₂

5. If any electrons are left over, place them on the central atom. We will explain later that some atoms are able to accommodate more than eight electrons. 6. If the central atom has fewer electrons than an octet, use lone pairs from terminal atoms to form multiple (double or triple) bonds to the central atom to achieve an octet. This will not change the number of electrons on the terminal atoms. 7. Final check Always make sure all valence electrons are accounted for and that each atom has an octet of electrons, except for hydrogen (with two electrons). The central atom is usually the least electronegative element in the molecule or ion; hydrogen and the halogens are usually terminal. Now let's apply this procedure to some particular compounds, beginning with one we have already discussed. Write the Lewis Structure for H₂O. Solutions to Example 10.4.1 Steps for Writing Lewis Structures Example (PageIndex{1}) 1. Determine the total number of valence electrons in the molecule or ion. Each H atom (group 1) has 1 valence electron, and the O atom (group 16) has 6 valence electrons, for a total of 8 valence electrons. 2. Arrange the atoms to show specific connections. Because H atoms are almost always terminal, the arrangement within the molecule must be HOH. 3. Place a bonding pair of electrons between each pair of adjacent atoms to give a single bond. 4. Beginning with the terminal atoms, add enough electrons to each atom to give each atom an octet (two for hydrogen). Placing one bonding pair of electrons between the O atom and each H atom gives with 4 electrons left over. Each H atom has a full valence shell of 2 electrons. 5. If any electrons are left over, place them on the central atom. Adding the remaining 4 electrons to the oxygen (as two lone pairs) gives the following structure: 6. If the central atom has fewer electrons than an octet, use lone pairs from terminal atoms to form multiple (double or triple) bonds to the central atom to achieve an octet. Not necessary. 7. Final check. The Lewis structure gives oxygen an octet and each hydrogen 2 electrons. Write the Lewis structure for the (CH₂O) molecule Solutions to Example 10.4.2 Steps for Writing Lewis Structures Example (PageIndex{2}) 1. Determine the total number of valence electrons in the molecule or ion. Each hydrogen atom (group 1) has 1 valence electron, carbon (group 14) has 4 valence electrons, and oxygen (group 16) has 6 valence electrons, for a total of [(2)(1) + 4 + 6] = 12 valence electrons. 2. Arrange the atoms to show specific connections. Because carbon is less electronegative than oxygen and hydrogen is normally terminal, C must be the central atom. 3. Place a bonding pair of electrons between each pair of adjacent atoms to give a single bond. Placing a bonding pair of electrons between each pair of bonded atoms gives the following; 6 electrons are used, and 6 are left over. 4. Beginning with the terminal atoms, add enough electrons to each atom to give each atom an octet (two for hydrogen). Adding all 6 remaining electrons to oxygen (as three lone pairs) gives the following: Although oxygen now has an octet and each hydrogen has 2 electrons, carbon has only 6 electrons. 5. If any electrons are left over, place them on the central atom. Not necessary. There are no electrons left to place on the central atom. 6. If the central atom has fewer electrons than an octet, use lone pairs from terminal atoms to form multiple (double or triple) bonds to the central atom to achieve an octet. To give carbon an octet of electrons, we use one of the lone pairs of electrons on oxygen to form a carbon-oxygen double bond: 7. Final check Both the oxygen and the carbon now have an octet of electrons, so this is an acceptable Lewis electron structure. The O has two bonding pairs and two lone pairs, and C has four bonding pairs. This is the structure of formaldehyde, which is used in embalming fluid. Write Lewis electron structures for CO₂ and SCl₂, a vile-smelling, unstable red liquid that is used in the manufacture of rubber. Answer CO₂. Answer SCl₂. The United States Supreme Court has the unenviable task of deciding what the law is. This responsibility can be a major challenge when there is no clear principle involved or where there is a new situation not encountered before. Chemistry faces the same challenge in extending basic concepts to fit a new situation. Drawing of Lewis structures for polyatomic ions uses the same approach, but tweaks the process a little to fit a somewhat different set of circumstances. Recall that a polyatomic ion is a group of atoms that are covalently bonded together and which carry an overall electrical charge. The ammonium ion, (ce{NH4^+}), is formed when a hydrogen ion (left(ce{H^+}) right)) attaches to the lone pair of an ammonia (left(ce{NH3}) right)) molecule in a coordinate covalent bond. Figure (PageIndex{3}). The ammonium ion. (CK12 License) When drawing the Lewis structure of a polyatomic ion, the charge of the ion is reflected in the number of total valence electrons in the structure. In the case of the ammonium ion: (1 \times \text{ce{N}}) atom (5 valence electrons \times 4 \text{ \times } \text{ce{H}}) atoms (4 \times 1 = 4) valence electrons subtract 1 electron for the (1+) charge of the ion total of 8 valence electrons in the ion It is customary to put the Lewis structure of a polyatomic ion into a large set of brackets, with the charge of the ion as a superscript outside of the brackets. Draw the Lewis electron dot structure for the sulfate ion. Answer (CK12 License) As important and useful as the octet rule is in chemical bonding, there are some well-known violations. This does not mean that the octet rule is useless—quite the contrary. As with many rules, there are exceptions, or violations. There are three violations to the octet rule. Odd-electron molecules represent the first violation to the octet rule. Although they are few, some stable compounds have an odd number of electrons in their valence shells. With an odd number of electrons, at least one atom in the molecule will have to violate the octet rule. Examples of stable odd-electron molecules are NO, NO₂, and ClO₂. The Lewis electron dot diagram for NO is as follows: Although the O atom has an octet of electrons, the N atom has only seven electrons in its valence shell. Although NO is a stable compound, it is very chemically reactive, as are most other odd-electron compounds. Electron-deficient molecules represent the second violation to the octet rule. These stable compounds have less than eight electrons around an atom in the molecule. The most common examples are the covalent compounds of beryllium and boron. For example, beryllium can form two covalent bonds, resulting in only four electrons in its valence shell: Boron commonly makes only three covalent bonds, resulting in only six valence electrons around the B atom. A well-known example is BF₃. The third violation to the octet rule is found in those compounds with more than eight electrons assigned to their valence shell. These are called expanded valence shell molecules. Such compounds are formed only by central atoms in the third row of the periodic table or beyond that have empty d orbitals in their valence shells that can participate in covalent bonding. One such compound is PF₅. The only reasonable Lewis electron dot diagram for this compound has the P atom making five covalent bonds: Formally, the P atom has 10 electrons in its valence shell. Identify each violation to the octet rule by drawing a Lewis electron dot diagram. a. With one Cl atom and one O atom, this molecule has 6 + 7 = 13 valence electrons, so it is an odd-electron molecule. A Lewis electron dot diagram for this molecule is as follows: Identify the violation to the octet rule in XeF₂ by drawing a Lewis electron dot diagram. Answer The Xe atom has an expanded valence shell with more than eight electrons around it. Lewis dot symbols provide a simple rationalization of why elements form compounds with the observed stoichiometries. A plot of the overall energy of a covalent bond as a function of internuclear distance is identical to a plot of an ionic pair because both result from attractive and repulsive forces between charged entities. In Lewis electron structures, we encounter bonding pairs, which are shared by two atoms, and lone pairs, which are not shared between atoms. Lewis structures for polyatomic ions follow the same rules as those for other covalent compounds. There are three violations to the octet rule: odd-electron molecules, electron-deficient molecules, and expanded valence shell molecules.