Chemistry 20

Lesson 9 – Lewis Diagrams

Lessons 9 to 14 provides us with a discussion about Chemical Bonding, specifically covalent bonding and ionic bonding. We will learn how to predict the formulas and shapes of different molecules, and we will be introduced to the properties of matter that result in differing boiling and melting points.

I. Atomic Orbitals

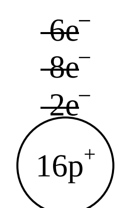
In Lesson 2 we learned about the modified Bohr Model of the atom as a simple but useful model to explain why atoms form bonds with each other. Recall that a chemical bond **involves the electrons only** – the nuclei of the different atoms are not affected. Further, the reason atoms interact, according to quantum theory, is that all atoms attempt to meet two criteria:

1. To be **electrically neutral** (# of protons = # of electrons).
2. To acquire the more **stable electron structure of the <u>nearest</u> noble gas**.

Chemical bonds are formed because most atoms can not satisfy these two criteria individually. To accomplish both criteria, atoms form **bonds** with one another. Two types of bonds are possible:

1. **An exchange of electrons** – where one atom gives up electrons to another atom. In other words, one atom loses electrons to become a cation while the other atom gains the electrons to become an anion. The loss and gain of electrons results in an **ionic bond** between the atoms. We will discuss this type of bond in more detail in Lesson 14.
2. **Sharing of electrons** – where atoms share electrons with one another. This type of bond is referred to as **molecular (covalent) bonding**. This type of bonding is described in greater detail in this lesson and in Lessons 10 to 12.

We will now take our general understanding and go into greater detail about the sharing of electrons. Using sulphur as an example, in our previous work we learned that sulphur has 16 electrons and they are distributed as shown in the diagram to the right. Since the two inner shells of electrons are full and complete they do not participate in bonding. The outer or **valence** shell has six electrons. However, according to the modern theory of atoms (i.e. quantum theory), the electrons in a shell do not exist together in one orbit. Rather, electron shells of an atom consist of **orbitals**. The rules that describe orbitals are:



sulphur
S

1. An orbital of any atom may be occupied by one or two electrons, but not more than two.
2. Since atoms can have up to eight electrons in the outer shell (except hydrogen and helium) there are **four** orbitals with a maximum of two electrons per orbital.
3. In order to have the same electron structure as the nearest noble gas atoms require each orbital to have two electrons for a total of eight electrons. This is referred to as having a full **octet** of valence electrons.

4. The first four valence electrons around a nucleus will, to minimise repulsion, occupy one orbital each. When more than four electrons are in the valence level, they must double up in the orbitals. (This will be made clearer in the next section on Lewis diagrams.)

5. The electrons in singly occupied orbitals are available for sharing with other atoms and are known as **bonding electrons**. Electrons that are paired in an orbital are not available for bonding. These nonbonding electron pairs are known as **lone pairs**.

The above rules for the electron occupancy of orbitals are purposely limited to those atoms which have four valence orbitals (with the exception of hydrogen and helium). These rules will cover the atoms considered in the Chemistry 20 bonding unit. There are many molecules and compounds that require a broader and more comprehensive theory to explain their structures that we will not consider in this course.

II. Lewis Diagrams for Atoms

The **octet rule** and the concepts of **bonding electrons** and **lone pair electrons** are applied in Lewis (electron-dot) diagrams. These diagrams provide a simple and convenient means for keeping track of the distribution of valence electrons involved in the **covalent bonding** of atoms. Lewis diagrams for chemical bonding were developed in 1916 by G. N. Lewis in order to describe covalent bonding.

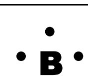
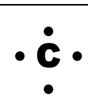
Lewis diagrams use the atomic symbol of the atom to represent both the nucleus and the filled innermost energy levels of the atom which do not participate in bonding. In addition, dots (or crosses) are arranged around the atomic symbol to represent the valence electrons. Unpaired electrons are separated from one another and paired electrons are shown close together. Thus •**Y**• indicates two unpaired (bonding) electrons, whereas $\ddot{\mathbf{Y}}$ indicates a lone electron pair.

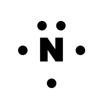
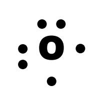
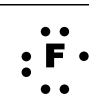
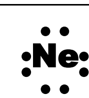
The rules for placing electrons in the four available orbitals of an atom are:

1. One electron dot is placed in each of the four available orbitals before any electron pairing occurs.

2. If there are five to eight valence electrons, a second electron dot is drawn into the singly occupied orbitals.

To illustrate this idea, consider the Lewis diagrams for the elements in period 2 from lithium to neon.

Atom	# of valence electrons	# of orbitals	Lewis Diagram	# of lone pairs	# of bonding electrons
Li	1	4		0	1
Be	2	4		0	2
B	3	4		0	3
C	4	4		0	4

Atom	# of valence electrons	# of orbitals	Lewis Diagram	# of lone pairs	# of bonding electrons
N	5	4		1	3
O	6	4		2	2
F	7	4		3	1
Ne	8	4		4	0

Note that the maximum number of bonding electrons for an atom is 4. Both carbon and silicon have four bonding sites which accounts for the vast number of carbon-based and silicon-based compounds found in nature. On the other hand, noble gases like neon have no bonding electrons which accounts for their inert, non-reactive nature.

(In this unit Lewis diagrams will be limited to atoms and molecules obeying the octet rule with the exception of hydrogen. Lewis diagrams are useful for situations beyond the octet rule but are not studied in high school chemistry.)

III. Assignment – part A

Complete the following table:

Atom	# Valence e^-	# Orbitals	Lewis Diagram	# Lone pairs	# Bonding e^-
S					
Si					
P					
Cl					
Br					
Be					

IV. Covalent bonding and Lewis diagrams for molecules

Again, the basic idea is that atoms will exchange or share electrons with other atoms in order to achieve a noble-gas-like electron structure (i.e., a full/stable octet). The **sharing** of electrons to achieve stable octets results in the formation of molecules held together by **covalent bonding**. The theory of covalent bonding involves the following:

1. **Covalent** bonding occurs between **non-metallic atoms**.
2. Atoms seek to fill their valence orbitals (i.e. two electrons in each of the four valence orbitals (with the exception of hydrogen)).
3. Unpaired electrons in the valence orbitals are available for bonding (i.e., sharing with other atoms).

4 Two atoms containing unpaired electrons can form a covalent bond by sharing the previously unpaired electrons.

5. The term **covalent** refers to the **sharing** of electrons between non-metallic atoms. The bonding is explained in terms of the **simultaneous attraction** of the pair of shared electrons by two nuclei.

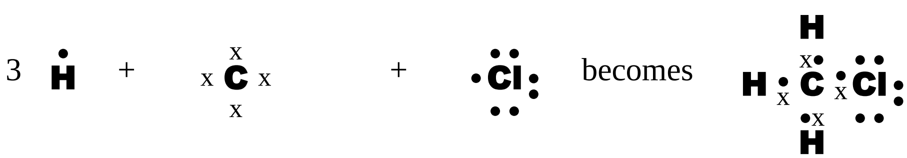
6. The number of covalent bonds that each atom forms (the **bonding capacity**) is limited. Bonding capacity is determined by the tendency of the atom to share electrons until it has the same electron structure as that of the nearest noble gas. This especially stable structure usually consists of eight electrons in the valence energy level and is called the **octet** structure.

Consistent with the octet rule, atoms share an appropriate number of electrons to attain a stable octet configuration characteristic of the noble gases. (Hydrogen, an important exception to the octet rule, attains two valence electrons which is the structure characteristic of helium.) Further, we draw Lewis diagrams to illustrate how covalent bonding results in the formation of molecules. The steps for drawing Lewis diagrams for molecules are:

1. Draw the Lewis diagrams for each of the atoms in the molecule.
2. Unpaired electrons (called **bonding electrons**) are available for sharing to form a covalent bond.
3. Paired electrons (called **lone pairs**) do not partake in the bonding.
4. The atom with the most bonding electrons (called the central atom) is placed in the centre with the other atoms bonded to it.
5. In the resulting Lewis diagram all electrons must be paired and each atom, except hydrogen, must be surrounded by an octet of electrons.

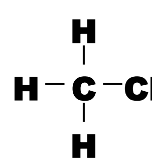
Example 1

Chloromethane has the following chemical formula: CH_3Cl Draw the Lewis diagram for this molecule.



The use of dots and crosses to represent valence electrons does not imply that there are different kinds of electrons. All electrons are identical. The use of dots and crosses is merely used as an aid to keep track of which electrons originated from which atoms.

Bonding representation is often simplified by omitting the lone electron pairs and by substituting a dash (–) for each bonding pair of electrons. The resulting representation is called a **structural formula**. The structural formula for chloromethane is shown to the right. Note that the structural formula does not generally represent the shape of the molecule, it merely represents which atoms are bonded together.



V. Assignment – part B

Molecular Substance	Molecular Formula	Lewis Diagrams of each atom	Lewis diagram of Molecule	Structural diagram
chloroform	$CHCl_3$			
ammonia	NH_3			
water	H_2O			
bromine	Br_2			
oxygen difluoride	OF_2			
fluoromethane	CH_3F			
chlorine	Cl_2			
nitrogen triodide	NI_3			
carbon tetrafluoride	CF_4			
dibromo-chloromethane	$CHBr_2Cl$			
silicon tetrafluoride	SiF_4			
hydrogen chloride	HCl			
methanol	CH_3OH			
ethanol	C_2H_5OH			
hydroxide	OH^-			