Chapter 4 : Chemical Bonds

Learning Outcomes:

a) To understand the bond in molecules

b) To understand the basic principles of bonds

c) To understand the ionic compounds and covalent compounds

Essential Vocabulary:

Ionic bond, Born-Haber Cycle, Covalent bond, Multiple Bond, Naming ionic compound, and Naming Covalent compounds

Overview and Introduction:

Molecules form three types of bonds: ionic, covalent, and metallic. This chapter delves deeply into ionic and covalent bonds, covering both basic principles and detailed steps for determining each bond's properties. The naming of ionic and covalent compounds is thoroughly discussed, with examples provided.

I: Chemical Bonds

As you are aware, the chemical formula for water is H2O, while salt is NaCl. Every day, we use a lot of water and salt. It is critical to understand how water is formed from two hydrogens (2H) and one oxygen (O), and salt from sodium (Na) and chloride (Cl). The combination of hydrogen and oxygen, as well as sodium and chlorine, has a strong attraction to form bonds. Because of the strong attraction between atoms, many bonds form, including ionic bonds, covalent bonds, and metallic bonds. This chapter discusses the fundamental concepts and properties of these bonds.

I-1: Ionic Bonds

As you know, I've enjoyed discussing salt in class as an example of something we all use on a daily basis, so I want you to be more engaged in class. To form a bond between Na and Cl, the first Na must be positively ionized by losing an electron, as shown in Figure 1a, producing sodium ion (Na^+). As shown in Figure 1a, one electron is in the outside orbital, so sodium loses an electron to comply with the octet rule and becomes a sodium ion. In contrast to Na, Cl gains one electron from sodium before forming the chloride ion (Cl^-), as illustrated in Figure 1b. Because the outer orbital contains seven electrons, chloride prefers to gain one electron in order to satisfy the octet rule, as illustrated in figure 1b. Metals in groups I and II of the periodic table lose electrons and form positive ions. Nonmetal atoms in groups V, VI, and VII prefer to gain electrons and produce negative ions. Positive metal ions are strongly electrostatically attracted to nonmetal negative ions, resulting in the formation of ionic bonds.

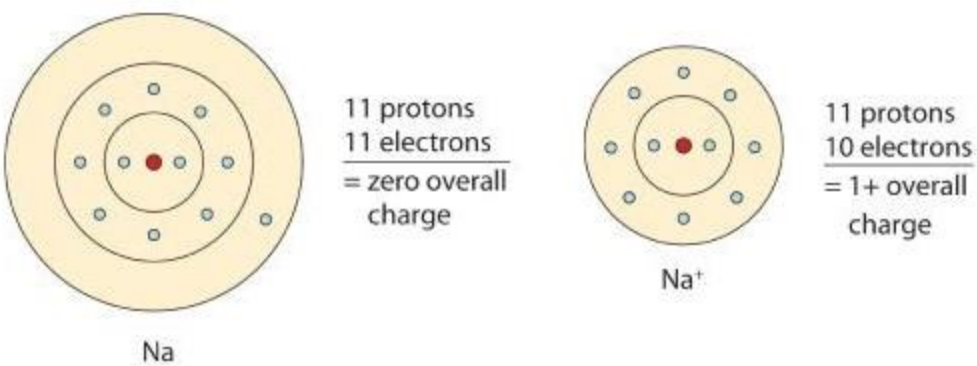


Figure 1a: The Formation of a Sodium Ion. On the left, a sodium atom has 11 electrons. On the right, the sodium ion only has 10 electrons and a 1^+ charge. Neutral sodium atom on left has 11

protons and 11 electrons. Sodium ion on right has 11 protons and 10 electrons, with a 1^+ overall charge.

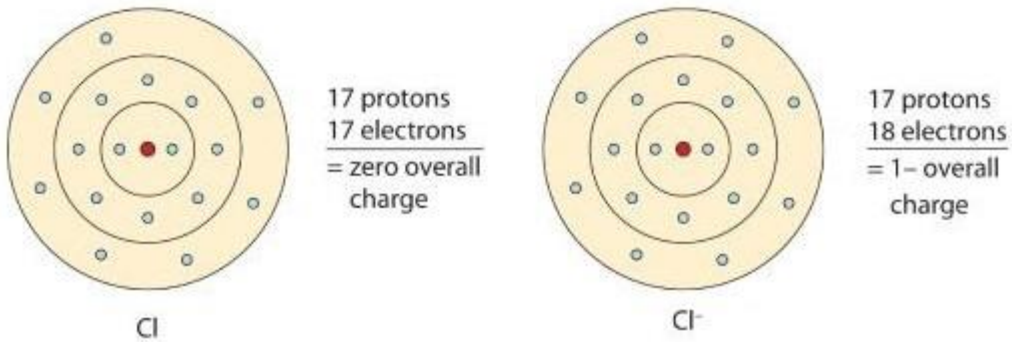


Figure 1b: The Formation of a Chlorine Ion. On the left, the chlorine atom has 17 electrons. On the right, the chloride ion has 18 electrons and has a 1^- charge. Neutral chlorine atom on left has 17 protons and 17 electrons. Sodium ion on right has 17 protons and 18 electrons, with a 1^- overall charge.

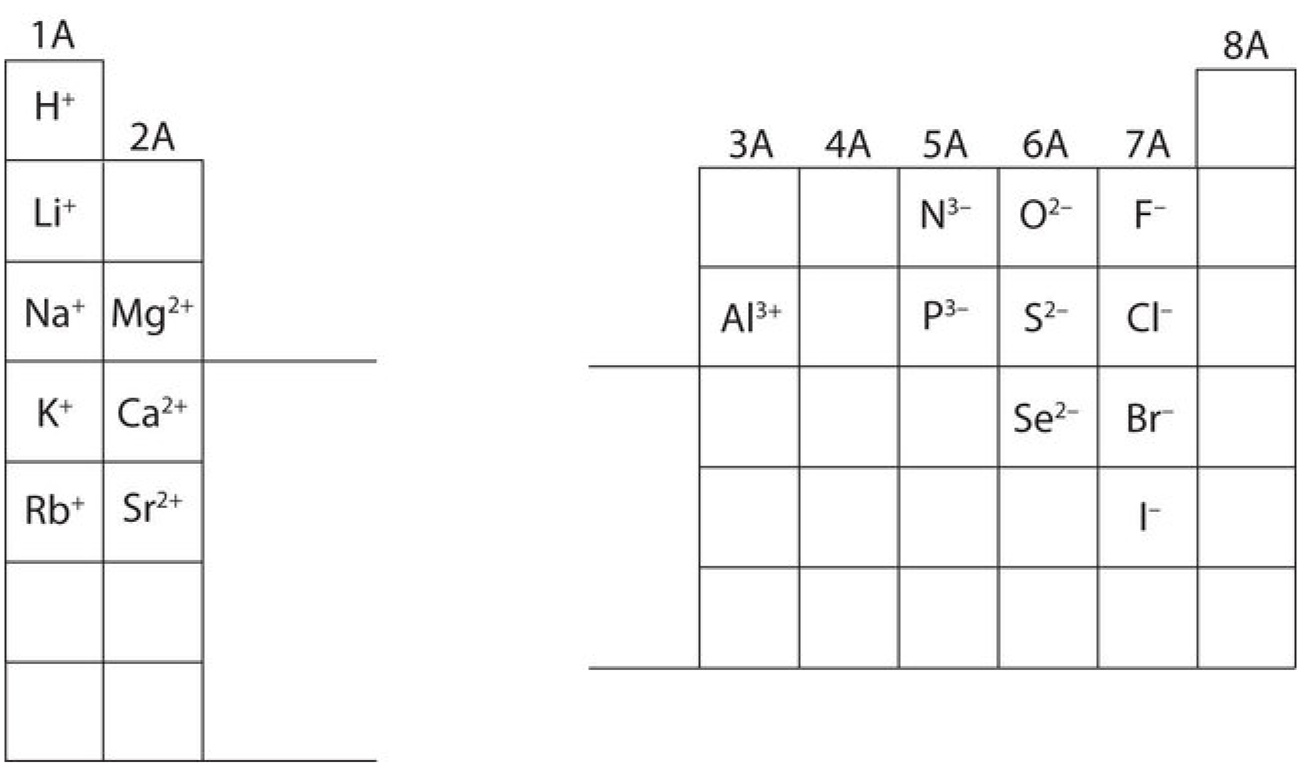


Figure 1c: Predicting Ionic Charges. The charge that an atom acquires when it becomes an ion is related to the structure of the periodic table. Within a group (family) of elements, atoms form ions of a certain charge.

https://chem.libretexts.org/Courses/Eastern_Mennonite_University/EMU%3A_CHEM_155_-_Matter_and_Energy_(Yoder)/Unit_1/1%3A_Matter/1.12%3A_Ions_-_Losing_and_Gaining_Electrons

To meet the zero charge of compounds, NaCl is formed by combining one Na^+ and one Cl^-. Figure 2 depicts how the Na^+ and Cl^- ions are orderly arranged in the compounds. Na and Cl have equal coordination numbers because NaCl is a one-to-one compound. The larger green ions represent Cl^-, while the smaller purple ions represent Na^+.

Figure 3 depicts the cubic crystal structure of NaCl, which is transparent, hard, and colorless in the solid state.

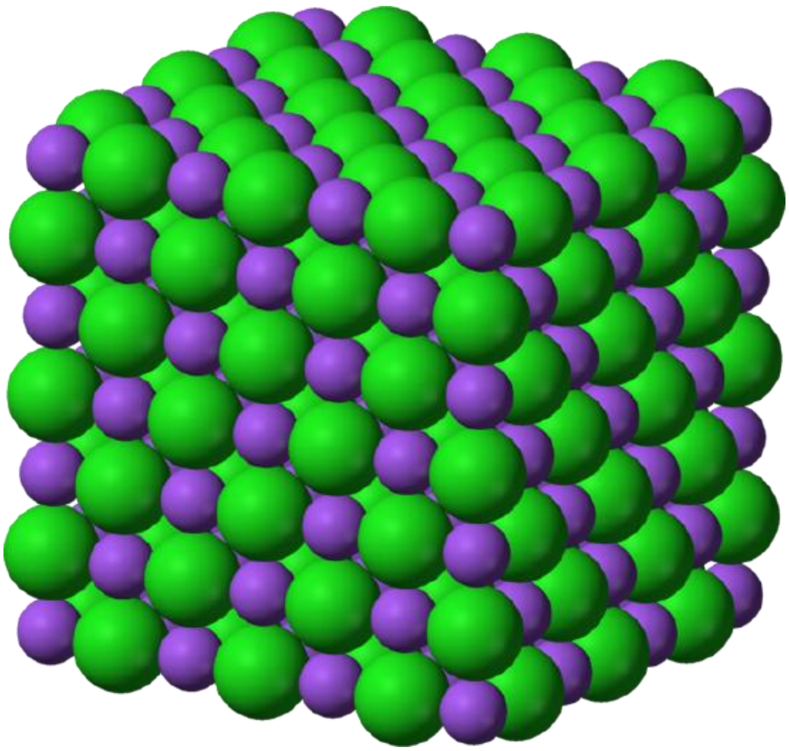


Figure 2: The crystal structure of sodium chloride, NaCl, a typical ionic compound. The smaller purple spheres represent sodium cations, Na^+, and the larger green spheres represent chloride anions, $Cl^−$. (Public Domain; Benjah-bmm27 via Wikipedia)

Figure 3: Crystal of Sodium chloride, also known as salt or halite, is an ionic compound with the chemical formula NaCl, representing a 1:1 ratio of sodium and chloride ions.

https://physicsopenlab.org/2018/01/22/sodium-chloride-nacl-crystal/

Ionic compounds are hard to deform and melt at high temperatures. For example, NaCl has a melting point of 801°C. These properties are the result of the regular arrangement of ions in the crystalline lattice, as well as the strong electrostatic attractive forces between ions with opposing charge. These electrostatic attractive forces in ionic compounds can be approximated by lattice energy, which can be calculated using the Born-Haber equation. ΔH_{lat} of NaCl has a lattice energy of 788 kJ/mol. Lattice energy (LE) is also called lattice enthalpy. The lattice energy is the amount of energy needed to convert one mole of a solid ionic compound into gaseous ions. The reaction is $NaCl(s) \rightarrow Na+(g) + Cl-(g)$. Figure 4 depicts all of the mechanisms.

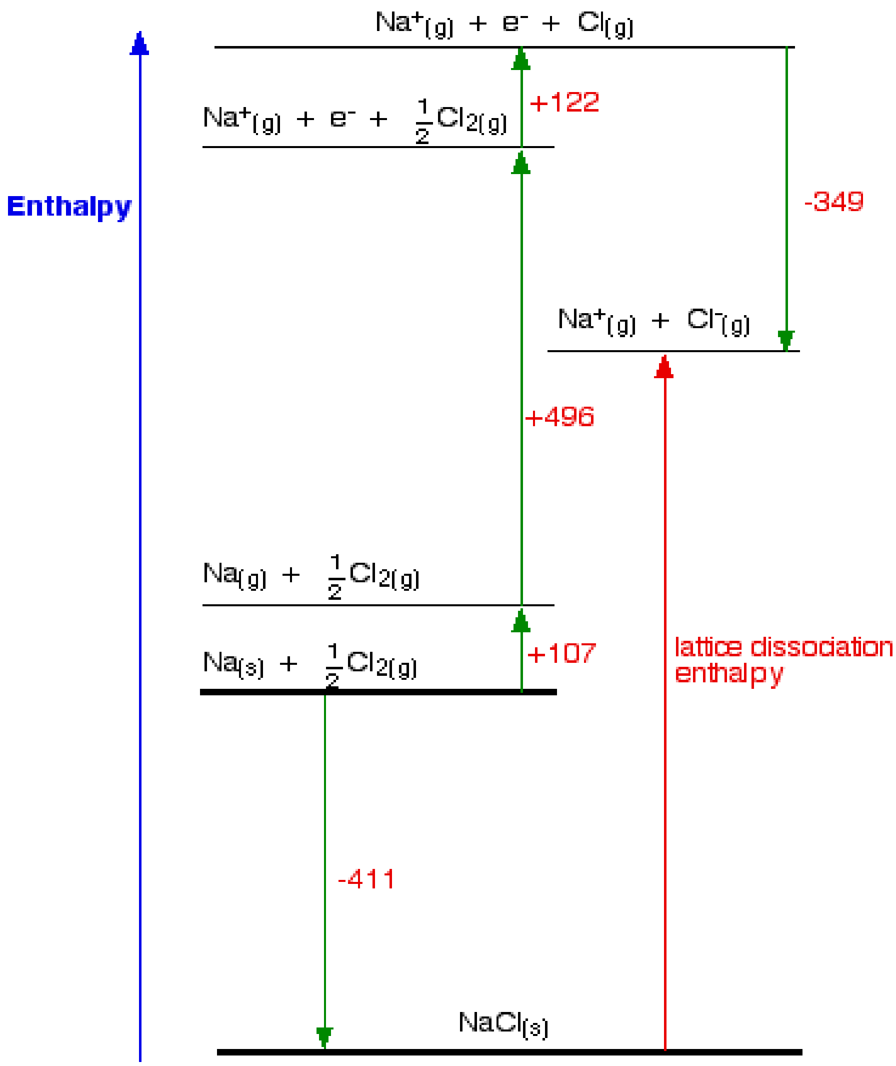


Figure 4: The Born-Haber cycle of Formation of NaCl.

https://chem.libretexts.org/Bookshelves/Inorganic_Chemistry/Supplemental_Modules_and_Websites_%28Inorganic_Chemistry%29/Crystal_Lattices/Thermodynamics_of_Lattices/Lattice_Enthalpies_and_Born_Haber_Cycles

The Born-Haber cycle of formation of NaCl is consisted by six major steps in formation of NaCl.

For the lattice enthalpy (LE) can be calculated as following:

The standard enthalpy of formation: -411

The atomization enthalpy of sodium: +107

The first ionization energy of sodium: +496

The atomization enthalpy of chlorine: +122

The first electron affinity of chlorine: -349

And finally, we have the positive and negative gaseous ions that we can convert into the solid sodium chloride using the lattice formation enthalpy.

-411 + LE = +107 + 496 + 122 - 349

LE = +107 + 496 + 122 - 349 + 411

LE = +787 kJ mol^{-1}

Let us look at another example, the CaF_2 ionic compound. As you recall, the compound (NaCl) is made up of sodium and chlorine in equal parts. However, in ionic compounds, the Ca/F ratio is one (Ca) to two (F). Figure 1C illustrates that Ca belongs to Group II and loses two electrons to become Ca 2^+, whereas F belongs to Group VII and gains one electron to become F^{-1}. As a result, one Ca 2^+ binds to two F^{-1} to form the compound CaF_2, which has a total charge of zero.

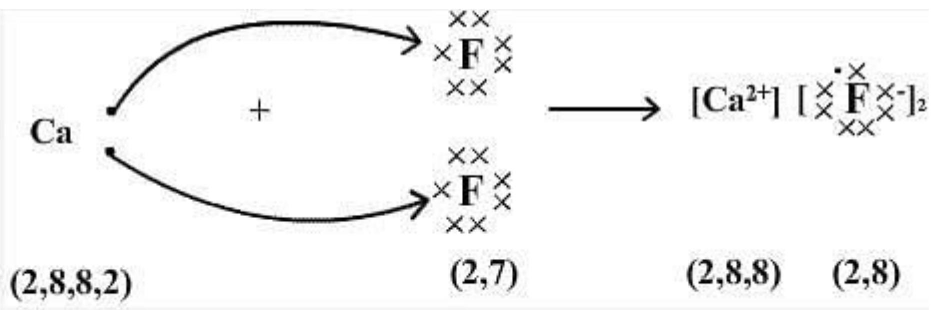


Table 1: Common Cations of Metals That Form More Than One Ion

Cation	Systematic Name	Common Name	Cation	Systematic Name	Common Name
Cr^{2+}	chromium(II)	chromous	Cu^{2+}	copper(II)	cupric
Cr^{3+}	chromium(III)	chromic	Cu^{+}	copper(I)	cuprous
Mn^{2+}	manganese(II)	manganous*	Hg^{2+}	mercury(II)	mercuric
Mn^{3+}	manganese(III)	manganic*	Hg_2^{2+}	mercury(I)	mercurous†
Fe^{2+}	iron(II)	ferrous	Sn^{4+}	tin(IV)	stannic
Fe^{3+}	iron(III)	ferric	Sn^{2+}	tin(II)	stannous
Co^{2+}	cobalt(II)	cobaltous*	Pb^{4+}	lead(IV)	plumbic*
Co^{3+}	cobalt(III)	cobaltic*	Pb^{2+}	lead(II)	plumbous*

Table 1 shows the list of metals which have more than one positive charges. the common polyatomic ions are listed in the table 2. The various metal ions and polyatomic ions also can use to form the ionic compounds with the specific combinations.

Table 2: Common Polyatomic Ions and Their Names

Formula	Name of Ion
NH_4^+	ammonium
$CH_3NH_3^+$	methylammonium
OH^-	hydroxide
O_2^{2-}	peroxide
CN^-	cyanide
SCN^-	thiocyanate
NO_2^-	nitrite
NO_3^-	nitrate
CO_3^{2-}	carbonate
HCO_3^-	hydrogen carbonate, or bicarbonate
SO_3^{2-}	sulfite
SO_4^{2-}	sulfate
HSO_4^-	hydrogen sulfate, or bisulfate
PO_4^{3-}	phosphate
HPO_4^{2-}	hydrogen phosphate
$H_2PO_4^-$	dihydrogen phosphate
ClO^-	hypochlorite
ClO_2^-	chlorite
ClO_3^-	chlorate
ClO_4^-	perchlorate
MnO_4^-	permanganate
CrO_4^{2-}	chromate

$Cr_2O_7^{2-}$	dichromate
$C_2O_4^{2-}$	oxalate
HCO_2^-	formate
$CH_3CO_2^-$	acetate
$C_6H_5CO_2^-$	benzoate

https://users.highland.edu/~jsullivan/genchem/s05_polyatomics

I-2: Covalent Bonding

All compounds do not have ionic compound properties because all atoms do not ionize and form ionic bonds. H_2, O_2, and H_2O are examples of simple molecules that form a covalent bond by sharing at least one pair of electrons rather than losing or gaining electrons. For example, a hydrogen molecule (H_2) has one electron in the outer orbital, while two hydrogen atoms share one electron in the valence orbital. As a result, two hydrogen atoms share two electrons, forming a single hydrogen molecule, as shown in Figures 1 and 2.

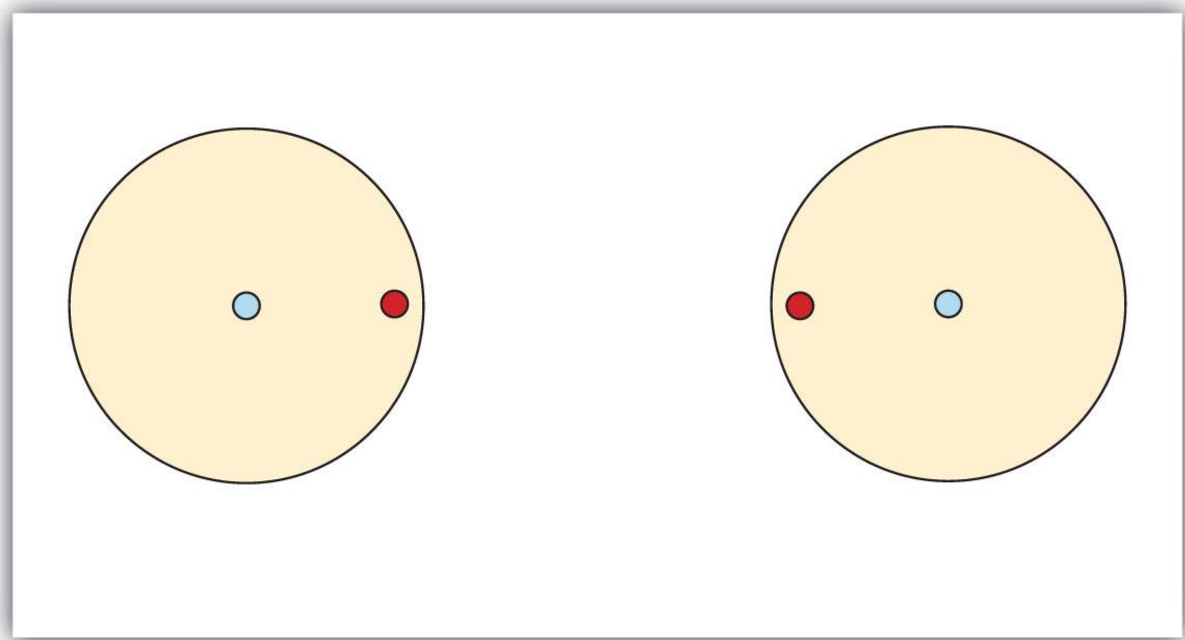


Figure 1: Individual Hydrogen Atoms

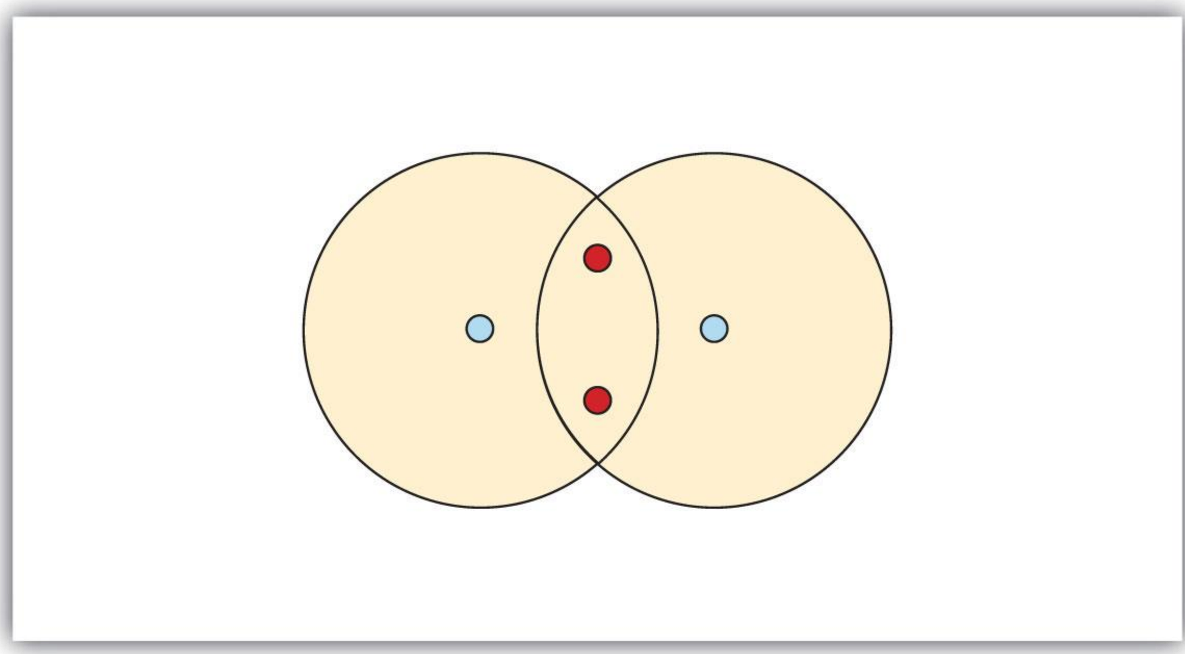


Figure 2: The Hydrogen Molecule

https://wou.edu/chemistry/courses/online-chemistry-textbooks/ch150-preparatory-chemistry/ch150-chapter-4-covalent-bonds-molecular-compounds/

Lewis electron dot diagrams are used to explain covalent bonds with the following steps: a- c.

a) Two Separate Hydrogen Atoms:

H•   •H

b) Hydrogen Molecule.

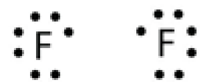
H:H

c) Hydrogen molecule is further simplified by using a dash.

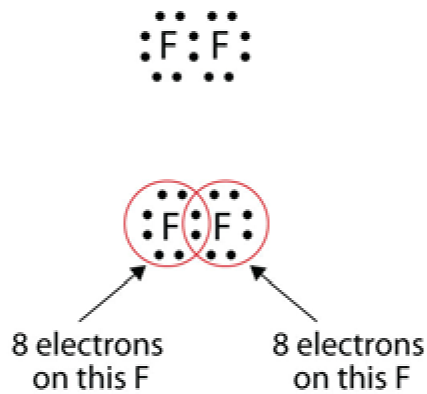
H—H

Another simple molecule F2 which has one covalent by sharing one electron from each F atom by the following steps;

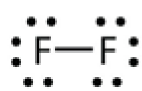
a) Two Separate Fluoride Atoms



b) Fluorine Molecule



c) Fluorine molecule is further simplified by using a dash line.



Two F share two electrons, which are known as bonding pairs, and electrons that are not shared by two F are known as lone pairs, as shown in step c. As we can see, H_2 and F_2 are diatomic gases that form covalent compounds due to their covalent bonds. Nonmetallic atoms and hydrogen combine to form new covalent compounds. The Lewis dot symbols for the common nonmetallic elements in Period 2 are listed in Table 3. The dots represent electrons in the valence orbital, and one can predict the number of bonds for each element based on the dots. For example, because C has four dots, it is likely that the carbon has four hydrogen bonds to meet the octet rule, which requires 8 electrons at the valence orbitals. As you can see, carbon shared 8 electrons with hydrogen, while each carbon shared two electrons with carbon.

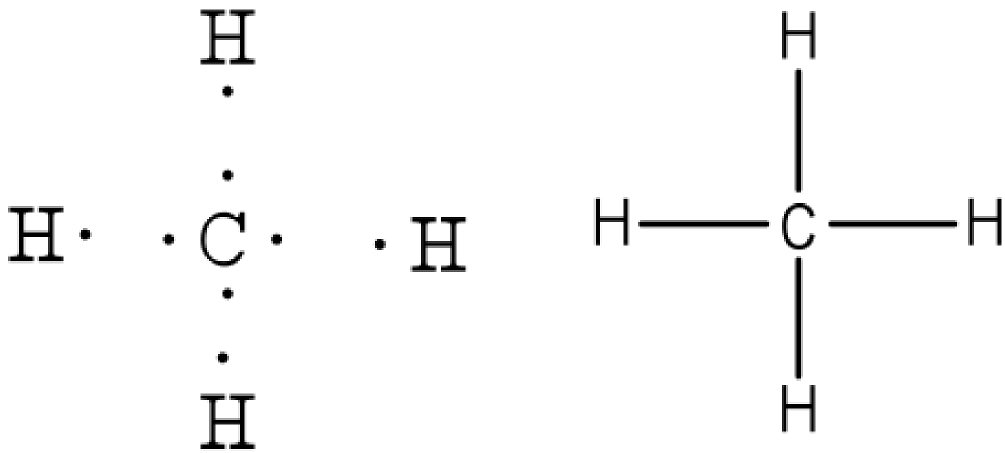
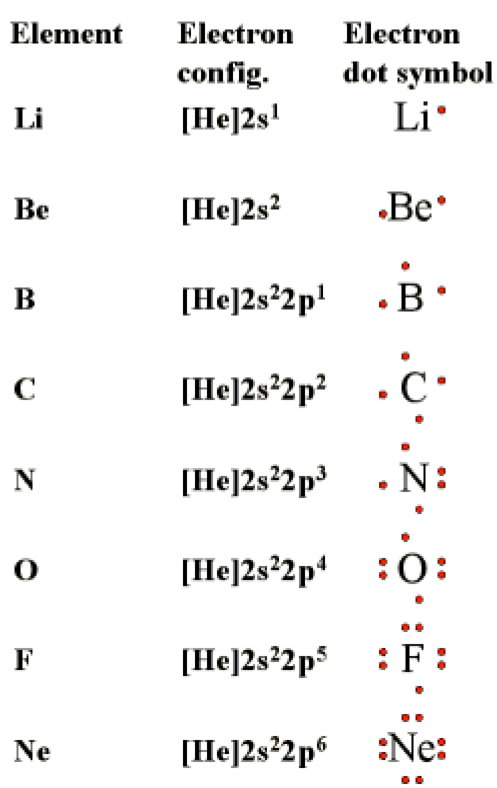


Figure 3: Lewis structure for methane (CH_4).



Element	Electron config.	Electron dot symbol
Li	[He]$2s^1$	Li
Be	[He]$2s^2$	Be
B	[He]$2s^2 2p^1$	B
C	[He]$2s^2 2p^2$	C
N	[He]$2s^2 2p^3$	N
O	[He]$2s^2 2p^4$	O
F	[He]$2s^2 2p^5$	F
Ne	[He]$2s^2 2p^6$	Ne

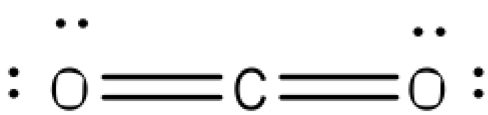
Table 3: Lewis Dot Symbols of the Elements in period 2

https://chem.libretexts.org/Courses/Valley_City_State_University/Chem_121/Chapter_8%3A_Chemical_Bonding_and_Molecular_Structures/8.1%3A_Lewis_Dot_Symbols_and_the_Octet_Rule

I-3: Multiple Bonds

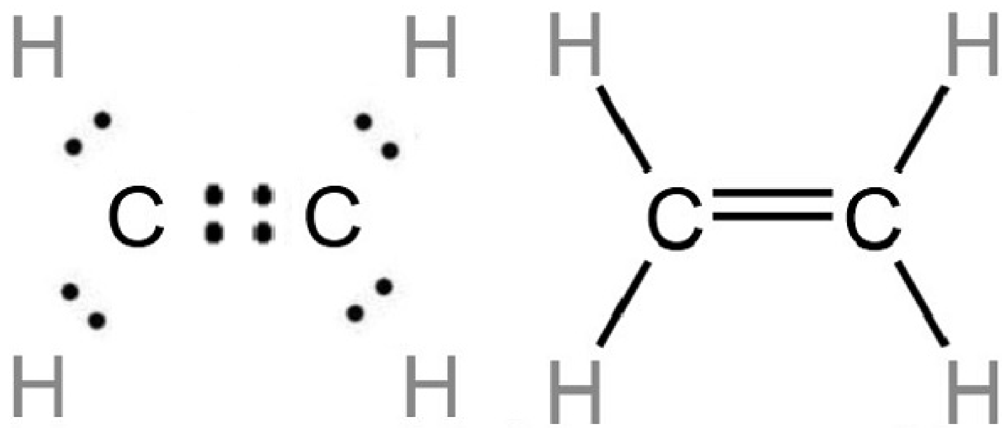
As we see the CH_4, H_2, F_2, only one pair electron is shared between atoms to make the covalent bond, this bond is called as single bond, as shown with a dash line. There are more covalent compounds with more than single bond such as double bond and triple bond. In CO_2, there are two bonding pairs between O and C are shred and makes double bond.

a) Separate Oxygen and Carbon



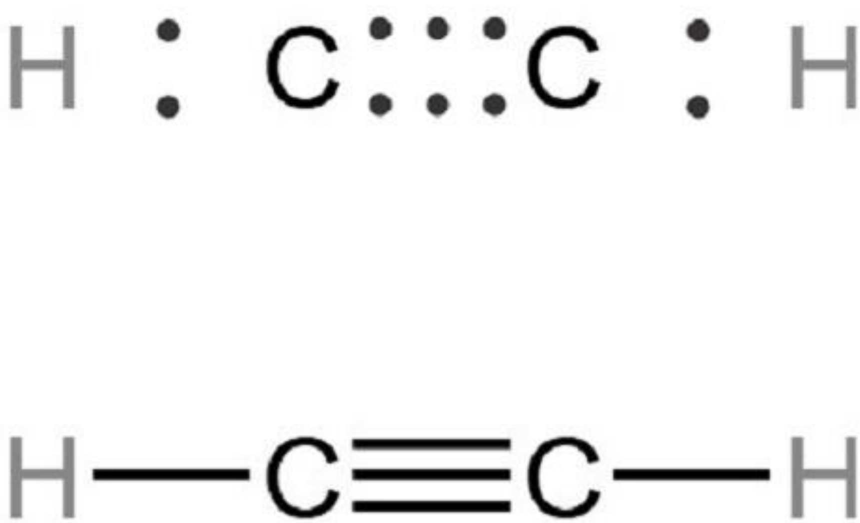
b) Lewis structure of CO_2

As you can see the Lewis structure of ethylene, there is a single bond (C-H) and double bond (C=C) in ethylene molecule.



c) Lewis structure of Ethylene

as you can see at the acetylene, six electrons are shared between Carbons, makes triple covalent bond.



d) Lewis structure of Acetylene

I-4: Bond Polarity

The explanation of the covalent bonds for H-H, F-F, C-H, and C-C states that since there is little variation in the electronegativity of the atoms in covalent bonds, the electrons are distributed equally among them. The electronegativity difference between atoms H (2.20) and C (2.55), as shown in table 4-5, is 0.35, which is less than 0. 5, indicating that the bond is non-polar. Additionally, there is no difference in electronegativity between the same elements, indicating that the covalent bond is non-polar. The electronegativity difference between atoms H (2.20) and Cl (3.15) in the covalent bond of HCl is 0.98, falling between 0.5 and 1.7, hence the term "polar covalent." The OH bond in water (H_2O) is polar covalent bond as well because the difference in electronegativity between H (2.20) and O (3.44) is 1.22, falling between 0.5 and 1.7. The bond between Na (0.94) and Cl (3.44) in salt, NaCl, has an electronegativity difference of 2.5, greater than 1.7, and is therefore referred to as an ionic bond as indicated at the table 4.

Table 4: The meaning of Absolute Differences in Electronegativity

Absolute Differences	Type of Bond Expected	Examples
1.7 or Greater	Ionic bond	NaCl
Between 0.5 to 1.7	Polar covalent bond	HCl, H_2O
0.5 or Less	Non-polar covalent bond	H_2, F_2, CH_4

Electronegativity of the Elements

1 H 2.20																	2 He no data
3 Li 0.98	4 Be 1.57											5 B 2.04	6 C 2.55	7 N 3.04	8 O 3.44	9 F 3.98	10 Ne no data
11 Na 0.93	12 Mg 1.31											13 Al 1.61	14 Si 1.90	15 P 2.19	16 S 2.58	17 Cl 3.16	18 Ar no data
19 K 0.82	20 Ca 1.00	21 Sc 1.36	22 Ti 1.54	23 V 1.63	24 Cr 1.66	25 Mn 1.55	26 Fe 1.83	27 Co 1.88	28 Ni 1.91	29 Cu 1.90	30 Zn 1.65	31 Ga 1.81	32 Ge 2.01	33 As 2.18	34 Se 2.55	35 Br 2.96	36 Kr 3.00
37 Rb 0.82	38 Sr 0.95	39 Y 1.22	40 Zr 1.33	41 Nb 1.6	42 Mo 2.16	43 Tc 1.9	44 Ru 2.2	45 Rh 2.28	46 Pd 2.20	47 Ag 1.93	48 Cd 1.69	49 In 1.78	50 Sn 1.96	51 Sb 2.05	52 Te 2.1	53 I 2.66	54 Xe 2.6
55 Cs 0.79	56 Ba 0.89	57-71	72 Hf 1.3	73 Ta 1.5	74 W 2.36	75 Re 1.9	76 Os 2.2	77 Ir 2.2	78 Pt 2.28	79 Au 2.54	80 Hg 2.00	81 Tl 1.62	82 Pb 2.33	83 Bi 2.02	84 Po 2.0	85 At 2.2	86 Rn no data
87 Fr 0.7	88 Ra 0.89	89-103	104 Rf no data	105 Db no data	106 Sg no data	107 Bh no data	108 Hs no data	109 Mt no data	110 Ds no data	111 Rg no data	112 Cn no data	113 Nh no data	114 Fl no data	115 Mc no data	116 Lv no data	117 Ts no data	118 Og no data

Low — High

57 La 1.10	58 Ce 1.12	59 Pr 1.13	60 Nd 1.14	61 Pm 1.13	62 Sm 1.17	63 Eu 1.2	64 Gd 1.2	65 Tb 1.22	66 Dy 1.23	67 Ho 1.24	68 Er 1.24	69 Tm 1.25	70 Yb 1.1	71 Lu 1.27
89 Ac 1.1	90 Th 1.3	91 Pa 1.5	92 U 1.38	93 Np 1.36	94 Pu 1.28	95 Am 1.3	96 Cm 1.3	97 Bk 1.3	98 Cf 1.3	99 Es 1.3	100 Fm 1.3	101 Md 1.3	102 No 1.3	103 Lr no data

Table 5: Periodic Table of the Elements: Electronegativity

https://www.google.com/

An essential factor in determining the bond in molecules is the electronegativity of the atom. "The comparative ability of atoms of an element to attract bonding electrons" in the bonding electron pairs is the definition of electronegativity. Figure 4a illustrates how the electron distribution of a nonpolar covalent bond is distributed equally among the atoms. As seen in figure 4b for the polar covalent bond, the electron in the HCl is more concentrated on the Cl side, making Cl partially negative charge and H partially positive charge. HCl is a dipole molecule as a result. The ionic bond in figure 4c is one in which the electron is completely transferred from Na to Cl as opposed to being shared.

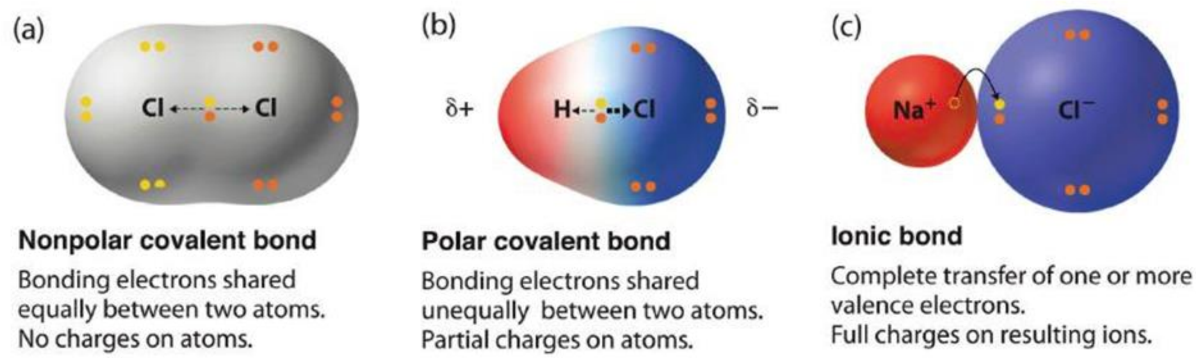


Figure 4. the electron distribution of non-polar covalent and polar covalent bonds

https://chem.libretexts.org/Courses/Smith_College/Organic_Chemistry_%28LibreTexts%29/02%3A_Polar_Covalent_Bonds_Acids_and_Bases/2.01%3A_Polar_Covalent_Bonds_-_Electronegativity

I-5: Names of Ionic Compounds

Metal and nonmetal components combine to form the majority of ionic compounds. As indicated in table 5, nonmetal elements with the suffix -ide come after metal names when naming binary ionic compounds that contain monatomic anions.

Table 5: Name of Ionic Compounds

Cation	Anion	Formula	Name
Na^{+}	Cl^{-}	NaCl	Sodium Chloride
Al^{3+}	Cl^{-}	$AlCl_3$	Aluminum chloride
Ca^{2+}	OH^{-}	$Ca(OH)_2$	Calcium hydroxide
Fe^{3+}	O^{2-}	Fe_2O_3	Iron (III) oxide

Suffix -ide is not necessary at the end of anion in ionic compounds containing polyatomic anions because table 6 illustrates that the anion's suffix is already present. For instance, sodium nitrate is known as $NaNO_3$. Ionic compounds containing transition metals, which typically have variable charges and monatomic anions; as table 7 illustrates, the metal names appear first with the Roman numeral for the charges in parenthesis, followed by the suffix-ide.

Table 6. Common Polyatomic Ions

Formula	Name
NH_4^+	Ammonium
OH^-	Hydroxide
NO_3^-	Nitrate
NO_2^-	Nitrite
$CH_3CO_2^-$	Acetate
CN^-	Cyanide
MnO_4^-	Permanganate
CO_3^{2-}	Carbonate
HCO_3^{--}	Bicarbonate
SO_3^{2-}	Sulfite
HSO_3^-	Bisulfite
SO_4^{2-}	Sulfate
PO_4^{3-}	Phosphate
HPO_4^{2-}	Hydrogen phosphate
$H_2PO_4^-$	Dihydrogen phosphate
SiO_3^{2-}	Silicate
CrO_4^{2-}	Chromate

Table 7. Names of Some Transition Metal Ionic Compounds

Transition Metal Ionic Compound	Name
$FeCl_3$	(III) chloride
Hg_2O	mercury(I) oxide
HgO	mercury(II) oxide
$Cu_3(PO_4)_2$	copper(II) phosphate

I-6: Covalent Compound Names

As indicated in tables 8 and 9, in order to name binary covalent compounds, the name of the first element comes first, followed by the stem of the other element with the prefix and suffix -ide. For instance, carbon dioxide is known as CO_2, and carbon monoxide is known as CO. When naming diatomic compounds, like oxygen and nitrogen, the prefix and suffix are not necessary; instead, the compounds are named using only their initial elements.

Table 8: Prefix for the Element

Number of Atoms in Compounds	Prefix for the element
1	mono-
2	di-
3	tri-
4	tetra-
5	penta-
6	hexa-
7	hepta-
8	octa-
9	nona-
10	deca-

Table 9: Stem Names of Common Elements

Element	Stem
Hydrogen	Hydr
Carbon	Carb
Nitrogen	Nitr
Oxygen	Ox
Fluorine	Fluor
Phosphorus	Phosph
Sulfur	Sulf
Chlorine	Chlor
Bromine	Brom
Iodine	Iod

Lab: Lewis Electron Dot Symbols

Outcomes: By the end of this section, you will be able to:

a) Write Lewis symbols for neutral atoms, ions and molecules

b) Understand the covalent bond

1) Introduction

All compounds do not have ionic compound properties because all atoms do not ionize and form ionic bonds. H_2, O_2, and H_2O are examples of simple molecules that form a covalent bond by sharing at least one pair of electrons rather than losing or gaining electrons. For example, a hydrogen molecule (H_2) has one electron in the outer orbital, while two hydrogen atoms share one electron in the valence orbital. As a result, two hydrogen atoms share two electrons, forming a single hydrogen molecule, as shown in Figures 1 and 2.

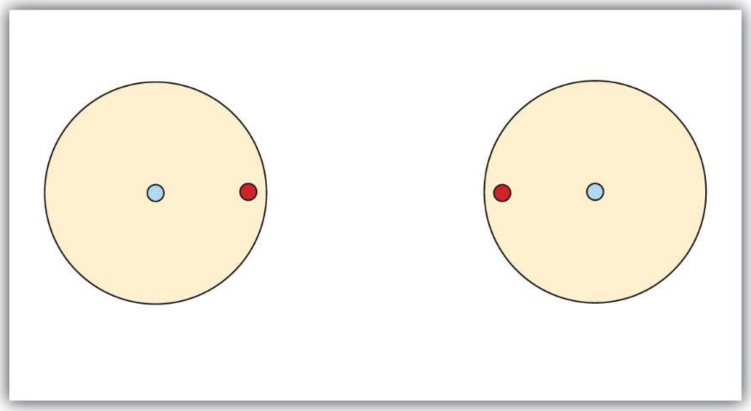


Figure 1: Individual Hydrogen Atoms

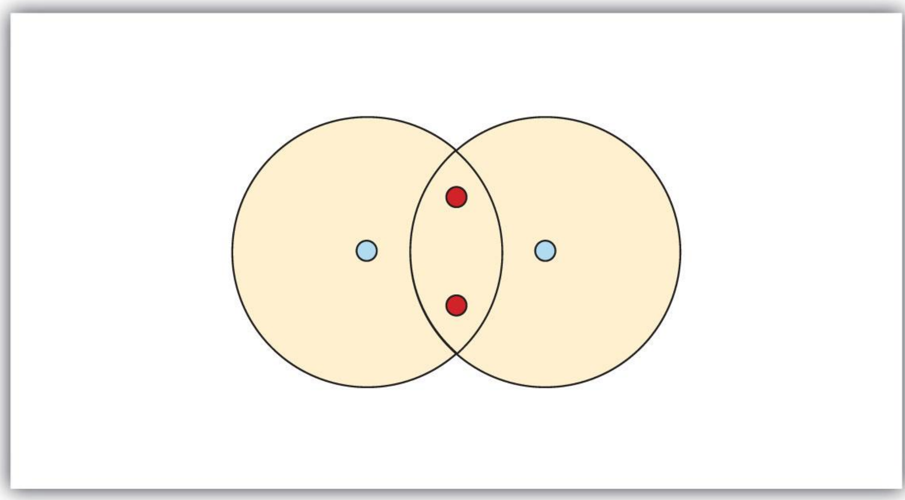


Figure 2: The Hydrogen Molecule

https://wou.edu/chemistry/courses/online-chemistry-textbooks/ch150-preparatory-chemistry/ch150-chapter-4-covalent-bonds-molecular-compounds/

Lewis electron dot diagrams are used to explain covalent bonds with the following steps: a- c.

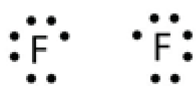
a) Two Separate Hydrogen Atoms:

b) Hydrogen Molecule.

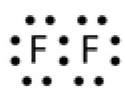
c) Hydrogen molecule is further simplified by using a dash.

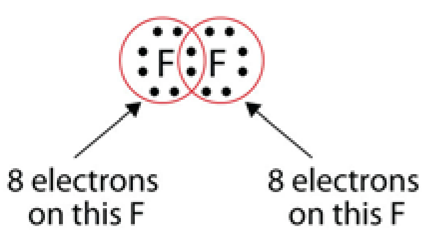
Another simple molecule F2 which has one covalent by sharing one electron from each F atom by the following steps;

a) Two Separate Fluoride Atoms



b) Fluorine Molecule





c) Fluorine molecule is further simplified by using a dash line.

Two F share two electrons, which are known as bonding pairs, and electrons that are not shared by two F are known as lone pairs, as shown in step c. As we can see, H_2 and F_2 are diatomic gases that form covalent compounds due to their covalent bonds. Nonmetallic atoms and hydrogen combine to form new covalent compounds. The Lewis dot symbols for the common nonmetallic elements in Period 2 are listed in Table 3. The dots represent electrons in the valence orbital, and one can predict the number of bonds for each element based on the dots. For example, because C has four dots, it is likely that the carbon has four hydrogen bonds to meet the octet rule, which requires 8 electrons at the valence orbitals. As you can see, carbon shared 8 electrons with hydrogen, while each carbon shared two electrons with carbon.

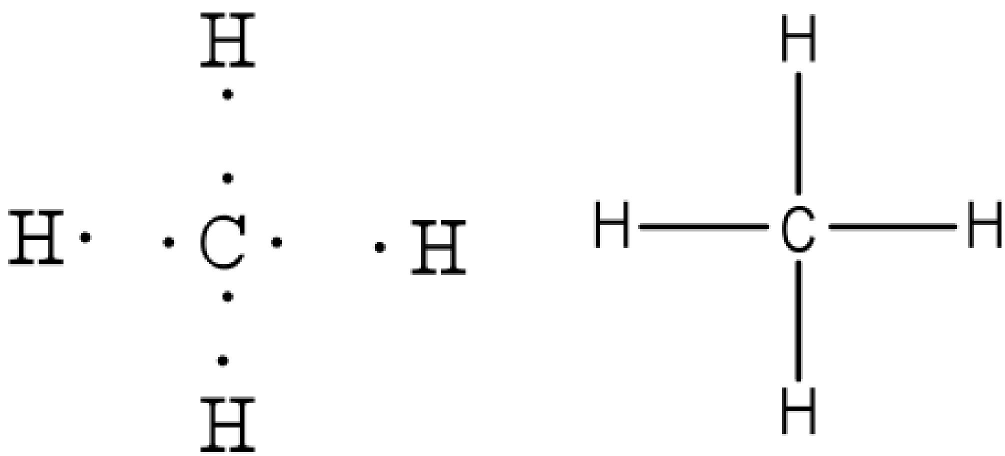
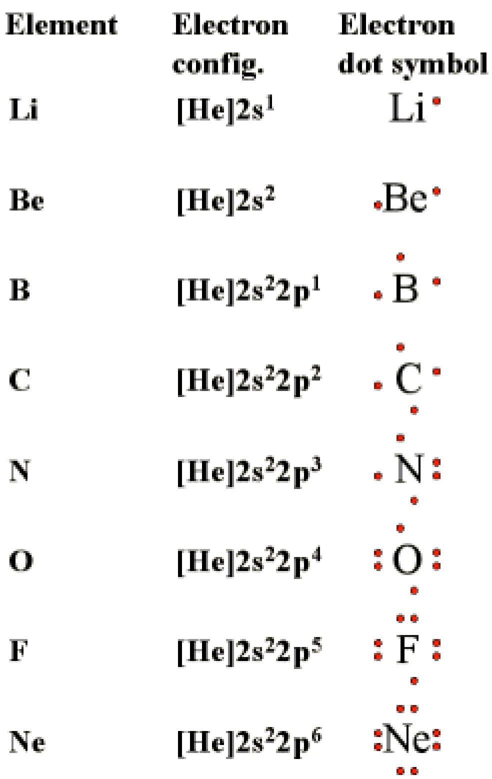


Figure 3: Lewis structure for methane (CH_4).



Element	Electron config.	Electron dot symbol
Li	[He]$2s^1$	Li
Be	[He]$2s^2$	Be
B	[He]$2s^2 2p^1$	B
C	[He]$2s^2 2p^2$	C
N	[He]$2s^2 2p^3$	N
O	[He]$2s^2 2p^4$	O
F	[He]$2s^2 2p^5$	F
Ne	[He]$2s^2 2p^6$	Ne

Table 3: Lewis Dot Symbols of the Elements in period 2

https://chem.libretexts.org/Courses/Valley_City_State_University/Chem_121/Chapter_8%3A_Chemical_Bonding_and_Molecular_Structures/8.1%3A_Lewis_Dot_Symbols_and_the_Octet_Rule

Questions

1) Draw the Lewis electron dot symbol for each atom

a) Boron

b) Nitrogen

c) Oxygen

d) Fluorine

2) Drew the Lewis structure for the following ions:

a) Hydronium ion ($H3O^+$)

b) Nitrate ion ($NO3^-$)

c) Bromide (Br^-)

3) Drew the Lewis structure for the following molecules

a) Carbon Dioxide (CO_2)

b) Water (H_2O)

c) Sodium Chloride (NaCl)