Ions, Noble Gases, Salts and Covalent Bonds

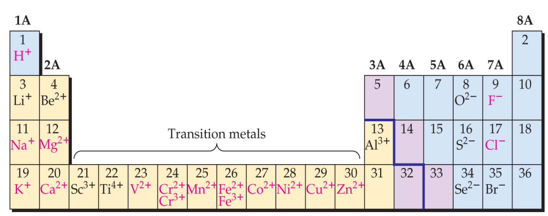
Where the Last Class Ended

Ions both positively charged and negatively charged are formed so that the electrons are arranged in filled shells.

This is predicted by Quantum Mechanics, and was observed when the periodic table was devised, which was long before Quantum Mechanics was developed.

Main Group Ions

- Group 1 elements tend to lose 1 electron
- Li^+, Na^+, K^+,Rb^+,Cs^+
- Group 2 elements lose 2 electrons
- Be^{2+},Mg^{2+}, Ca^{2+}, Sr^{2+}, Ba^{2+}
- Group 17 elements gain 1 e-
- Fl^-, Cl^-, Br^-, I^-
- Group 16 elements gain 2 e-
- O^{2-}, S^{2-},Se^{2-}, Te^{2-}



Ions formed by elements in the first four periods.

Naming Cations

Main group metal cations are named by identifying the metal, followed by the word "ion" as shown below:

- K^+ Potassium ion
- Mg^{2+} Magnesium ion
- Al^{3+} Aluminum ion

For transition metals which can form more than one type of cation, two systems are used.

Naming Anions

- Anions are named by replacing the ending of the element name with -ide, followed by the word "ion."

TABLE 4.3 Names of Some Common Anions

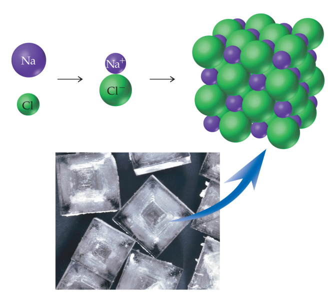
ELEMENT	SYMBOL	NAME
Bromine	Br^-	Bromide ion
Chlorine	Cl^-	Chloride ion
Fluorine	F^-	Fluoride ion
Iodine	I^-	Iodide ion
Oxygen	O^{2-}	Oxide ion
Sulfur	S^{2-}	Sulfide ion

Ionic Compounds

- NaCl (sodium chloride) and other salts are held together because of the attraction between the positive and negative ions.
- But we seldom find ionic compounds as single pairs of positive and negative charge.
- Ion pairs arranged in an alternating pattern in a lattice is much more stable than a single ion pair.

Na^+ and Cl^- ions in a sodium chloride crystal.

- Each Na^+ ion is surrounded by six Cl^- ions, and each Cl^- ion is surrounded by six Na^+ ions.
- The crystal is held together by ionic bonds—the attraction between oppositely charged ions.



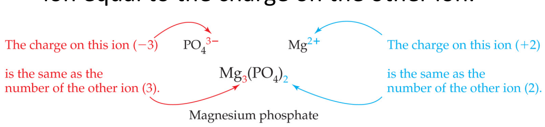
Properties of Ionic Compounds

- Ionic compounds are usually crystalline solids.
- Ions in an ionic solid are held rigidly in place by attraction to their neighbors and ions cannot move from one place in the lattice to another.
- Once an ionic solid is dissolved in water or melted, the ions can move freely and conduct electricity.
- Ionic compounds because the attractive forces present in the lattice are extremely strong. (Sodium chloride melts at 801ºC and boils at 1413ºC.)

- Ionic solids are not malleable and will shatter if struck sharply.
- Ionic compounds dissolve in water if the attraction between water and the ions overcomes the attraction of the ions for one another.
- Sodium chloride and some other familiar ionic compounds are very soluble and can be dissolved to make solutions of high concentration.
- Other ionic compounds are not water-soluble, because water is unable to overcome the ionic forces in many crystals.

4.10 Formulas of Ionic Compounds

- A chemical formula shows the simplest ratio of anions and cations required for a total charge of zero.
- A shortcut is to make the subscript of each ion equal to the charge on the other ion.



Once the numbers and kinds of ions in a compound are known, the formula is written using the following rules:

- List the cation first and the anion second; for example, NaCl not ClNa.
- Make sure to eliminate any common factors from the subscripts; for example, MgO not Mg_2O_2.
- Do not write the charges of the ions; for example, KF not K^+F^-
- Use parentheses around a polyatomic ion formula if it has a subscript; for example, $Al_2(SO_4)_3$ not Al_2SO_{43}.

Write the formulas of the following and name them.

- Mg and I
- Ca and Cl
- Na and I
- Ca and O
- Na and O
- Fe^{3+} and O
- Sn^{2+} and F
- Ti^{3+} and N

Covalent Bonds

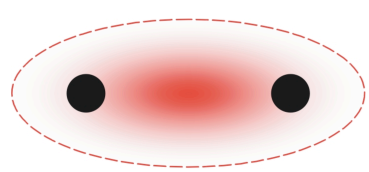
- All compounds are not salts but the elements that make up these compounds are still seeking the most stable state and for main group elements this means achieving a full orbital shell.
- Covalent bonds are formed when electrons are shared between atoms.
- Compounds formed by covalent bonds alone are called molecules.

Understanding Covalent Bonds

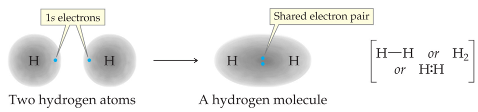
- From a Quantum Mechanic's view a molecule should be treated much like an atom, and all of the electrons on the molecule treated the same.
- This is not entirely practical, so chemists have come up with ways of using quantum mechanics to understand covalent bonds, without dispensing with the role of atoms.

Adding Up Atomic Orbitals to Make New Orbitals – Sigma Bonds

- Sigma Bonds Electrons are localized predominantly between the atoms

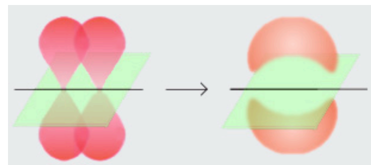


- Covalent bond formation in H_2 can be visualized by imagining that the two spherical 1*s* atomic orbitals blend together and overlap to give an egg-shaped molecular orbital.
- Both atoms share the two valence electrons and the stability of the closed shell electron configuration.
- The shared pair of electrons in a covalent bond can be represented as a line between atoms.



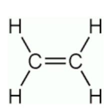
Adding Up Atomic Orbitals to Make New Orbitals-Pi Bonds

- Pi Bonds the orbitals are localized predominantly above and below the axis that goes between the two atoms.



Animation Showing Sigma and Pi Bonds

- On line Animation-C2H4
- http://www.youtube.com/watch?v=C2W-yDPcpl4



Molecular Orbitals

- As with atomic orbitals, molecular orbitals represent the probability of finding an electrons in a given region, not the location of an electron.
- An electron in a bonding orbital is not always between the atoms.
- The molecular orbitals are present whether or not electrons are in them.
- Whenever a bonding orbital is formed a nonbonding orbital is formed as well.

Electron Sharing Rules

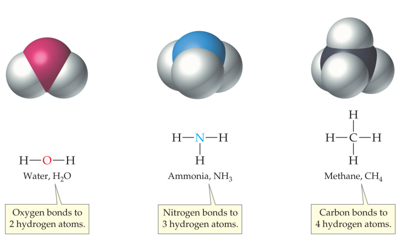
- Now we have considered that
- ✓ electrons in a molecule do not belong to a particular atom.
- ✓ the location of the electrons are not known
- ✓ the shared electrons are not always between the bonded atoms
- ✓ And countless other things

lets take a quick look at Lewis dot structures and electron counting.

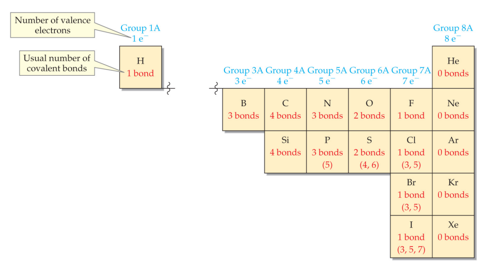
Electron Sharing Rules

- Water molecules consist of two hydrogen atoms joined by covalent bonds to an oxygen atom, ammonia molecules consist of three hydrogen atoms covalently bonded to a nitrogen atom, and methane molecules consist of four hydrogen atoms covalently bonded to a carbon atom.

Note that in all these examples, each atom shares enough electrons to achieve a noble gas configuration: each of the 2 hydrogens in the water molecule gains an electron, while the oxygen gains to electrons from the bonds if forms with each. (Sketch the dot structures of all 3.)

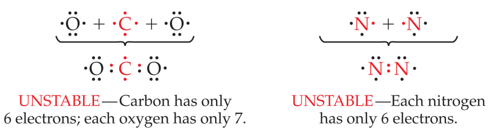


Number of bonds formed to achieve octet. Numbers in parentheses indicate possible numbers of bonds that result in exceptions to the octet rule.



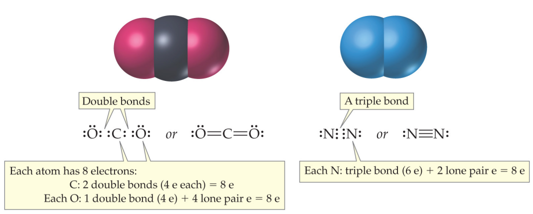
Multiple Covalent Bonds – When More Electrons Need to be Shared

- Molecules are not limited to sigma bonds and the need for pi bonds is apparent when making dot structures.

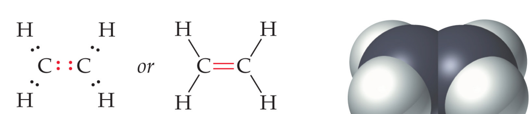


- **Single bond:** A covalent bond formed by sharing one electron pair. (1 sigma bond)
- **Double bond:** A covalent bond formed by sharing two electron pairs. (1 sigma and 1 pi bond)
- **Triple bond:** A covalent bond formed by sharing three electron pairs. (1 sigma and 2 pi bonds)
- Just as a single bond is represented by a single line between atoms, a double bond is represented by two lines between atoms and a triple bond by three lines.

Carbon, nitrogen, and oxygen are the elements most often present in multiple bonds.



- Ethylene, a simple compound used commercially to induce ripening in fruit, has the formula C_2H_4 .
- The only way for the 2 carbon atoms to have octets is for them to share 4 electrons in a double bond:

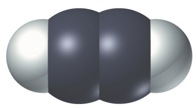


Ethylene—the carbon atoms share 4 electrons in a double bond.

- Acetylene, the gas used in welding, has the formula C_2H_2 .
- To achieve octets, two carbons share six electrons in a triple bond.
- In compounds with multiple bonds like ethylene and acetylene, each carbon atom still forms a total of four covalent bonds.

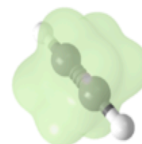
H : C ⋮⋮ C : H *or* H—C≡C—H

Acetylene—the carbon atoms share 6 electrons in a triple bond.



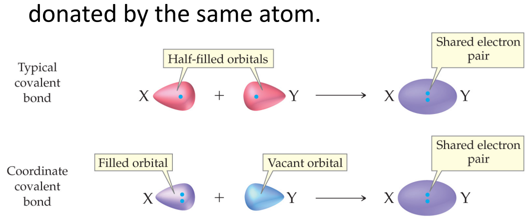
Acetylene

- This image shows electron distribution due to 2 pi orbitals on acetylene

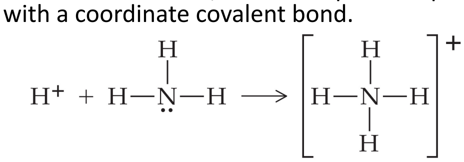


5.4 Coordinate Covalent Bonds

Coordinate covalent bond: The covalent bond that forms when both electrons are donated by the same atom.



- The ammonium ion, is an example of a species with a coordinate covalent bond.

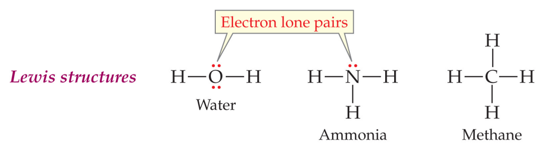


- Coordination compounds are an entire class of substances based on the ability of transition metals to form coordinate covalent bonds with nonmetals.
- Essential metal ions are held in enzyme molecules by coordinate covalent bonds.

5.5 Molecular Formulas and Lewis Structures

- **Molecular formula:** A formula that shows the numbers and kinds of atoms in one molecule of a compound.
- **Structural formula:** A molecular representation that shows the connections among atoms by using lines to represent covalent bonds.
- **Lewis structure:** A molecular representation that shows both the connections among atoms and the locations of lone-pair valence electrons.

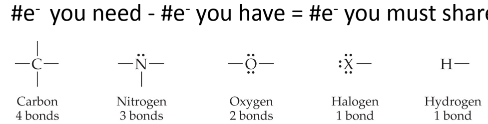
- ▶ The oxygen atom in H_2O shares 2 electron pairs with two hydrogen atoms and has 2 other pairs of valence electrons that are not shared in bonds.
- ▶ Such unshared pairs of valence electrons are called **lone pairs**.
- ▶ In NH_3, 3 electron pairs are used in bonds and there is 1 lone pair; in CH_4, all 4 pairs are bonding.



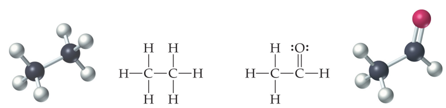
5.6 Drawing Lewis Structures

- H, C, N, O, and halogen atoms usually maintain consistent bonding patterns:
- H forms one covalent bond.
- C forms four covalent bonds.
- N forms three covalent bonds and has one lone pair of electrons.
- O forms two covalent bonds and has two lone pairs of electrons.
- Halogens form one covalent bond and have three lone pairs of electrons.

- Some Lewis structures can be assembled like a puzzle. Each piece an atom, fitting together by connecting bonding sites. Others require more hints.
- Find the number of e^- needed to satisfy each atom separately, $2e^-$ for H and $8e^-$ for all other atoms. Add up the number of valence electrons you actually have. If the octet rule is obeyed, then:
- $\#e^-$ you need - $\#e^-$ you have = $\#e^-$ you must share.



- Ethane is an example of a structure that can be made simply by connecting 2C's and 6H's at available bonding sites. For acetaldehyde, a hint helps.
- 4H's need 8e$^-$, 2C's and 1O need 24e$^-$,we need 32e$^-$. 4H's have 4e$^-$, 2C's have 8e$^-$, and 1O has 6e$^-$, we have 18e$^-$. (32e$^-$ - 18e$^-$ = 14e$^-$) It helps to know that 14e$^-$ must be shared in 7 bonds and the 4e$^-$ left over must form 2 unshared lone pairs.



Ethane, CH_3CH_3 Acetaldehyde, CH_3CHO

5.7 The Shapes of Molecules

- Molecular shapes can be predicted by noting how many bonds and electron pairs surround individual atoms and applying what is called the **valence-shell electron-pair repulsion (VSEPR)** model.
- The basic idea of the VSEPR model is that the negatively charged clouds of electrons in bonds and lone pairs repel each other, and keep as far apart as possible.

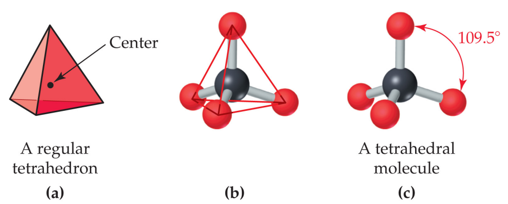
There are three step to applying the VSEPR model:

- Step 1: Draw a Lewis structure of the molecule, and identify the atom whose geometry is of interest.
- Step 2: Count the number of electron charge clouds surrounding the atom of interest.
- Step 3: Predict molecular shape by assuming that the charge clouds orient in space so that they are as far away from one another as possible.

TABLE 5.1 Molecular Geometry Around Atoms with 2, 3, and 4 Charge Clouds

NUMBER OF BONDS	NUMBER OF LONE PAIRS	NUMBER OF CHARGE CLOUDS	MOLECULAR GEOMETRY	EXAMPLE
2	0	2	Linear	O=C=O
3	0	3	Planar triangular	$H_2C{=}O$
2	1	3	Bent	O_2S
4	0	4	Tetrahedral	CH_4
3	1	4	Pyramidal	NH_3
2	2	4	Bent	H_2O

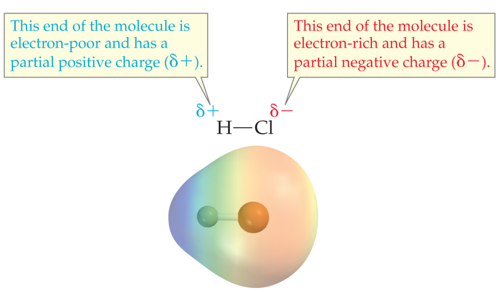
- Linear molecules have bond angles of 180°.
- Planar triangular molecules have bond angles of 120°.
- Tetrahedral molecules have bond angles of 109.5°.



5.8 Polar Covalent Bonds and Electronegativity

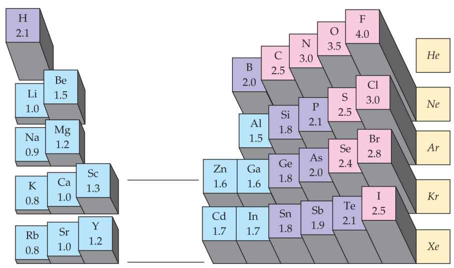
- Electrons in a covalent bond occupy the region between the bonded atoms.
- If the atoms are identical, as in H_2 and Cl_2, electrons are attracted equally to both atoms and are shared equally.
- If the atoms are not identical, however, as in HCl, the bonding electrons may be attracted more strongly by one atom than by the other and thus shared unequally. Such bonds are known as **polar covalent bonds**.

When charges separate in a neutral molecule, the molecule has a dipole moment and is said to be polar.

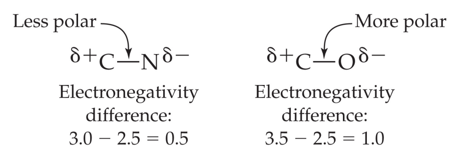


- In HCl, electrons spend more time near the chlorine than the hydrogen. Although the molecule is overall neutral, the chlorine is more negative than the hydrogen, resulting in partial charges on the atoms.
- Partial charges are represented by a δ- on the more negative atom and δ+ on the more positive atom.
- The ability of an atom to attract electrons is called the atom's **electronegativity**.
- Fluorine, the most electronegative element, assigned a value of 4, and less electronegative atoms assigned lower values.

Elements at the top right of the periodic table are most electronegative, those at the lower left are least electronegative. Noble gases are not assigned values.



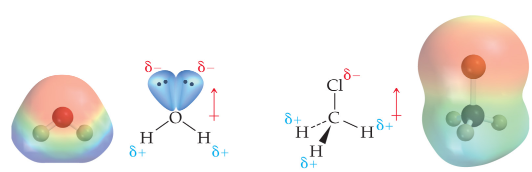
- As a rule of thumb, electronegativity differences of less than 0.5 result in nonpolar covalent bonds, differences up to 1.9 indicate increasingly polar covalent bonds, and differences of 2 or more indicate ionic bonds.
- There is no sharp dividing line between covalent and ionic bonds; most bonds fall somewhere in-between.



5.9 Polar Molecules

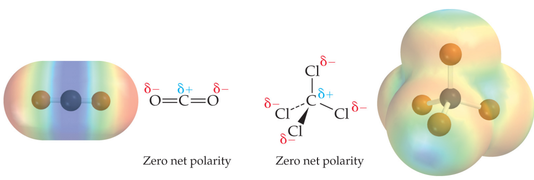
- Entire molecules can be polar if electrons are attracted more strongly to one part of the molecule than to another.
- Molecules polarity is due to the sum of all individual bond polarities and lone-pair contribution in the molecule.
- Polarity has a dramatic effect on the physical properties of molecules, particularly on melting points, boiling points, and solubility.

- Dipoles or polarity can be represented by an arrow pointing to the negative end of the molecule with a cross at the positive end resembling a + sign.



Water, H_2O Chloromethane, CH_3Cl

- Just because a molecule has polar covalent bonds does not mean that the molecule is polar overall.
- Carbon dioxide and tetrachloromethane molecules have no net polarity because their symmetrical shapes cause the individual bond polarities to cancel each other out.



Zero net polarity   Zero net polarity

5.10 Naming Binary Molecular Compounds

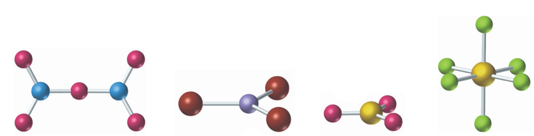
- The formulas of binary molecular compounds are written with the less electronegative element first.
- Name the first element in the formula, using a prefix to indicate the number of atoms.
- Name the second element in the formula, using the ending -ide as for anions along with a prefix if needed.

The prefix mono- is omitted for the first element.

TABLE 5.2 Numerical Prefixes Used in Chemical Names

NUMBER	PREFIX
1	mono-
2	di-
3	tri-
4	tetra-
5	penta-
6	hexa-
7	hepta-
8	octa-
9	nona-
10	deca-

The structures and names of several binary molecular compounds are shown below.



N_2O_5 *Di*nitrogen *pent*oxide BBr_3 Boron *tri*bromide SO_3 Sulfur *tri*oxide SF_6 Sulfur *hexa*fluoride

5.11 Characteristics of Molecular Compounds

TABLE 5.3 A Comparison of Ionic and Molecular Compounds

IONIC COMPOUNDS	MOLECULAR COMPOUNDS
Smallest components are ions	Smallest components are molecules
Usually composed of metals combined with nonmetals	Usually composed of nonmetals with nonmetals
Crystalline solids	Gases, liquids, or low-melting solids
High melting points	Low melting points
High boiling points (above 700°C)	Low boiling points
Conduct electricity when molten or dissolved in water	Do not conduct electricity
Many are water-soluble	Few are water-soluble
Not soluble in organic liquids	Many are soluble in organic liquids