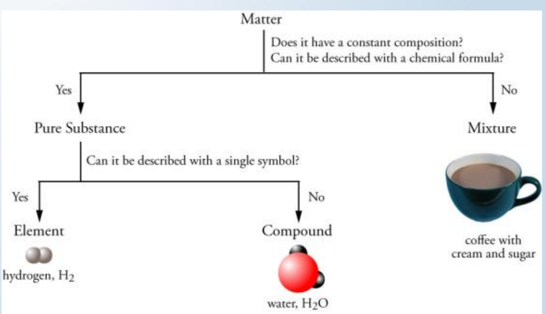
Chapter 5

Chemical Compounds

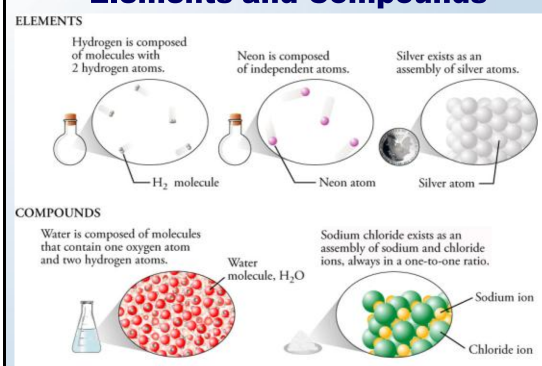
Elements, Compounds, and Mixtures

- **Element:** A substance that cannot be chemically converted into simpler substances; a substance in which all of the atoms have the same number of protons and therefore the same chemical characteristics.
- **Compound:** A substance that contains two or more elements, the atoms of these elements always combining in the same whole-number ratio.
- **Mixture:** A sample of matter that contains two or more pure substances (elements and compounds) and has variable composition.

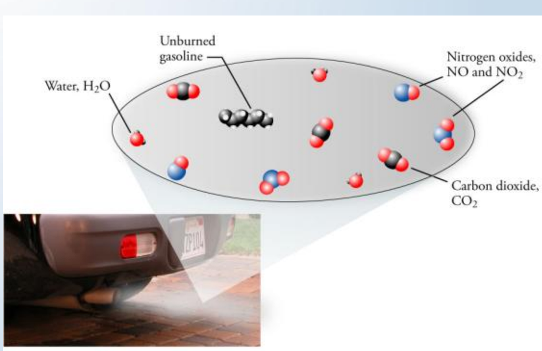
Classification of Matter



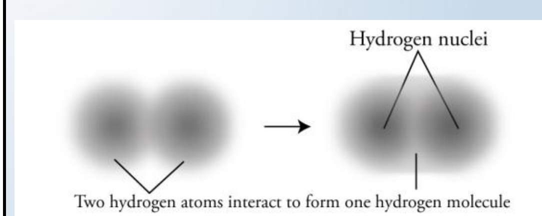
Elements and Compounds



Exhaust – a Mixture

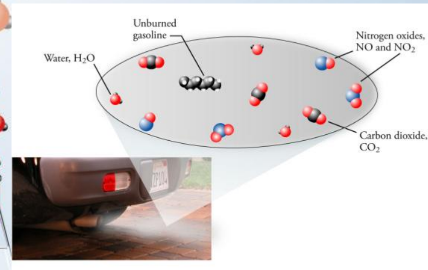


Covalent Bond Formation



Covalent Bond

- A link between atoms due to the sharing of two electrons.
- In compounds this bond forms between atoms of two or more nonmetallic elements.

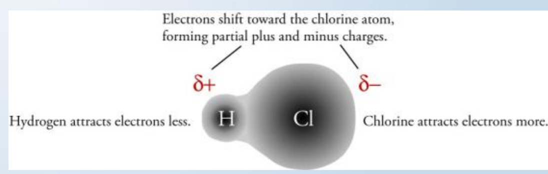


Covalent Bond

– If one atom in the bond attracts electrons more than the other atom, the electron negative charge shifts to that atom giving it a partial negative charge.

The other atom loses negative charge giving it a partial positive charge. The bond is called a ***polar covalent bond***.

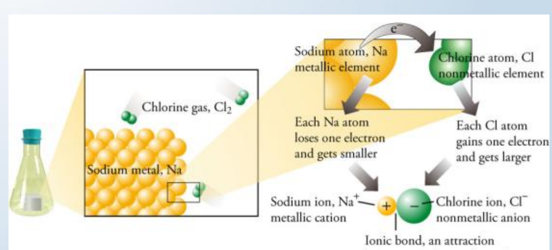
Polar Covalent Bond



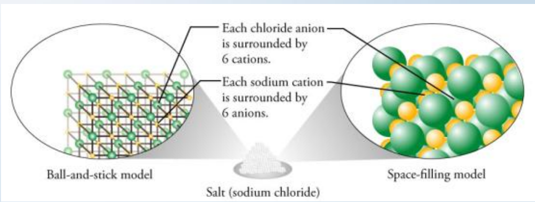
Ionic Bond

- The attraction between cation and anion.
- Atoms of nonmetallic elements often attract electrons so much more strongly than atoms of metallic elements that one or more electrons are transferred from the metallic atom (forming a positively charged particle or ***cation***), to the nonmetallic atom (forming a negatively charged particle or ***anion***).
- For example, an uncharged chlorine atom can pull one electron from an uncharged sodium atom, yielding Cl^- and Na^+.

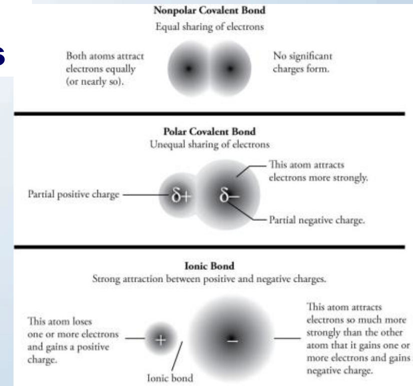
Ionic Bond Formation



Sodium Chloride, NaCl, Structure



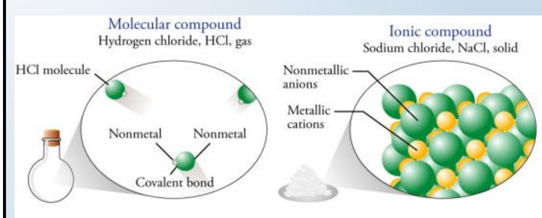
Bond Types



Types of Compounds

- All nonmetallic atoms usually leads to all covalent bonds, which from molecules. These compounds are called ***molecular compounds***.
- Metal-nonmetal combinations usually lead to ionic bonds and ***ionic compounds***.

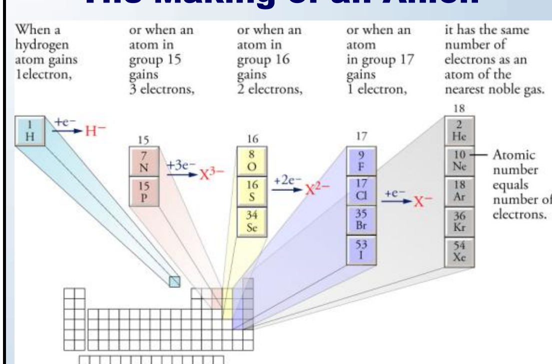
Classification of Compounds



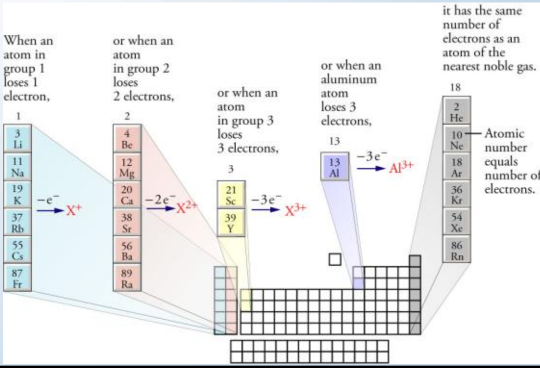
Summary

- **Nonmetal-nonmetal** combinations (e.g. HCl)
  – Covalent bonds
  – Molecules
  – Molecular Compound
- **Metal-nonmetal** combinations (e.g. NaCl)
  – Probably ionic bonds
  – Alternating cations and anions in crystal structure
  – Ionic compound

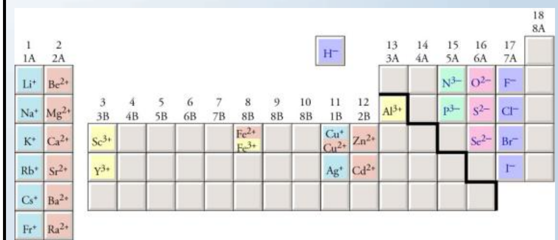
The Making of an Anion



The Making of a Cation



Monatomic Ions



Monatomic Ion Names

- Monatomic Cations
  - (name of metal)
    - Groups 1, 2, and 3 metals
    - Al^{3+}, Zn^{2+}, Cd^{2+}, Ag^+
  - (name of metal)(Roman numeral)
    - All metallic cations not mentioned above
- Monatomic Anions
  - (root of nonmetal name)ide

Monatomic Anions

Hydride H^-

Nitride N^{3-}

Phosphide P^{3-}

Oxide O^{2-}

Sulfide S^{2-}

selenide Se^{2-}

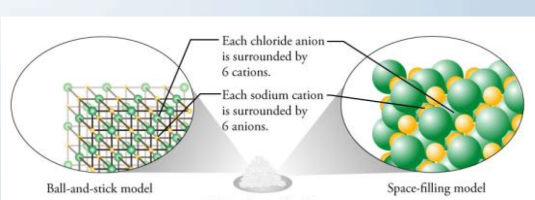
fluoride F^-

chloride Cl^-

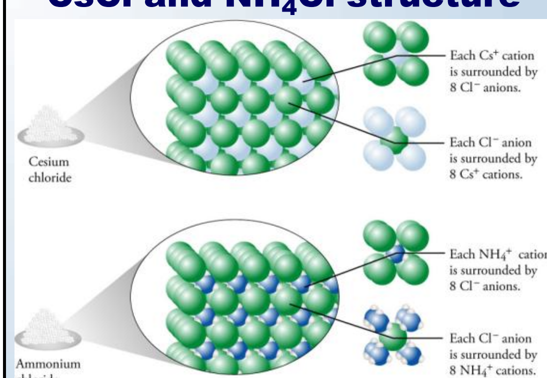
bromide Br^-

iodide I^-

Sodium Chloride, NaCl, Structure



CsCl and NH_4Cl structure



Models – Advantages and Disadvantages (1)

- They help us to *visualize*, *explain*, and *predict* chemical changes.

Models – Advantages and Disadvantages (2)

- One characteristic of models is that they ***change with time***.

Assumptions of the Valence-Bond Model

- Only the highest energy electrons participate in bonding (Group number)
- 2 electrons make a bond.

Valence Electrons

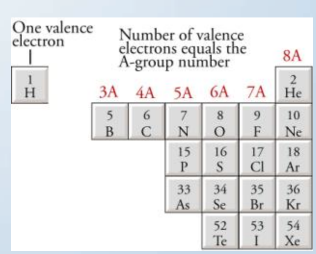
- The valence electrons for each atom are the most important electrons in the formation of chemical bonds.
- The number of valence electrons for the atoms of each element is equal to the element's A-group number on the periodic

Valence Electrons

- Covalent bonds often form to pair electrons and give the atoms of the elements eight valence electrons (an octet of valence electrons).
- Other than hydrogen and boron

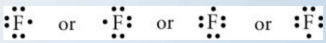
Valence Electrons

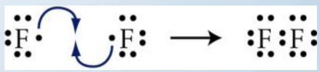
- ***Valence electrons*** are the highest-energy *s* and *p* electrons in an atom.



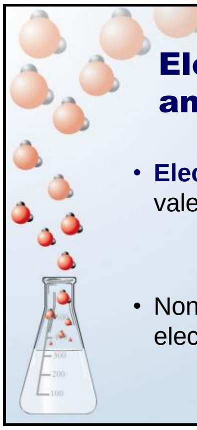
Fluorine

Valence electrons





Electron-Dot Symbols and Lewis Structures

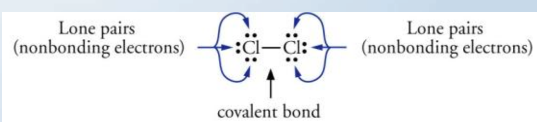


- **Electron-dot symbols** show valence electrons.

- Nonbonding pairs of valence electrons are called ***lone pairs***.

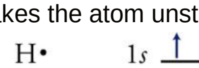
Lewis Structures

- ***Lewis structures*** represent molecules using element symbols, lines for bonds, and dots for lone pairs.

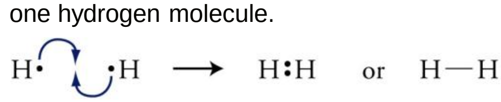


H_2 Formation

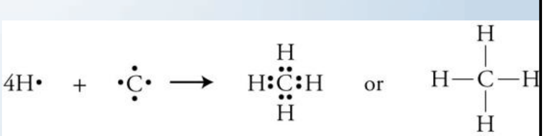
- The unpaired electron on a hydrogen atom makes the atom unstable.



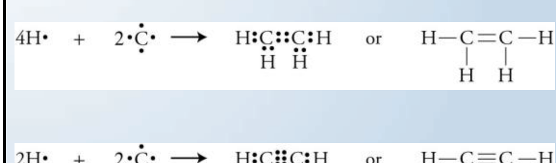
- Two hydrogen atoms combine to form one hydrogen molecule.



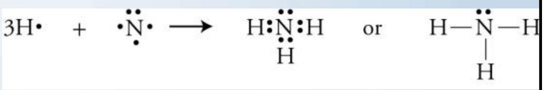
Carbon – 4 bonds



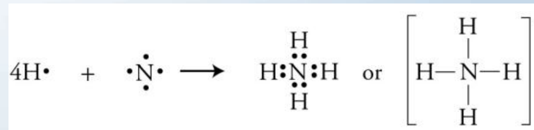
Carbon – Multiple Bonds



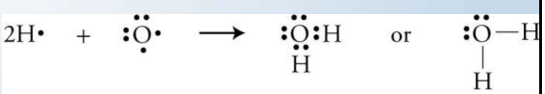
Nitrogen – 3 bonds & 1 lone pair



Nitrogen – 4 bonds

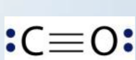


Oxygen – 2 bonds & 2 lone pairs

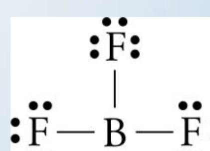


Oxygen – 1 bond & 3 lone pairs

Carbon – 3 bonds & 1 lone pair
Oxygen – 3 bonds & 1 lone pair



Boron – 3 bonds



Halogens – 1 bond & 3 lone pairs

Most Common Bonding Patterns for Nonmetals

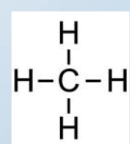
Element	# Bonds	# lone pairs
H	1	0
C	4	0
N, P	3	1
O, S, Se	2	2
F, Cl, Br, I	1	3

Drawing Lewis Structures

- This chapter describes a procedure that allows you to draw Lewis structures for many different molecules.
- Many Lewis structures can be drawn by attempting to give each atom in a molecule its most common bonding pattern.

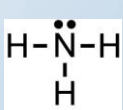
Lewis Structure for Methane, CH_4

- Carbon atoms usually have 4 bonds and no lone pairs.
- Hydrogen atoms have 1 bond and no lone pairs.



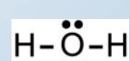
Lewis Structure for Ammonia, NH_3

- Nitrogen atoms usually have 3 bonds and 1 lone pair.
- Hydrogen atoms have 1 bond and no lone pairs.



Lewis Structure for Water, H_2O

- Oxygen atoms usually have 2 bonds and 2 lone pairs.
- Hydrogen atoms have 1 bond and no lone pairs.



Drawing Lewis Structures (1)

- **Step 1:** Determine the total number of valence electrons for the molecule or polyatomic ion.

Drawing Lewis Structures (2)

- **Step 2:** Draw a reasonable skeletal structure, using single bonds to join all the atoms.

Drawing Lewis Structures (3)

- **Step 3:** Subtract 2 electrons from the total for each of the single bonds (lines) described in Step 2.

Drawing Lewis Structures (4)

- **Step 4:** Try to distribute the remaining electrons as lone pairs to obtain a total of eight electrons around each atom except hydrogen and boron.

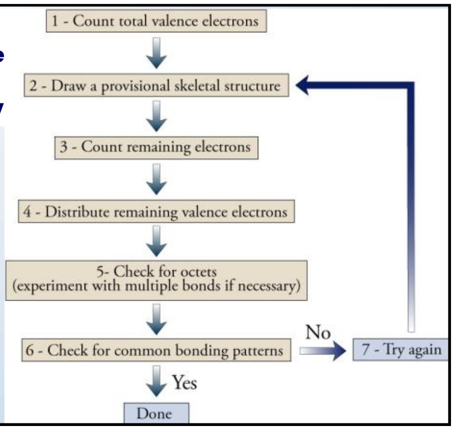
Drawing Lewis Structures (5)

- **Step 5:** Do one of the following.
  - If in Step 4 you were able to obtain an octet of electrons around each atom go to Step 6.
  - If you do not have enough electrons to obtain octets of electrons around each atom (other than hydrogen and boron), <u>convert one lone pair into a multiple bond for each two electrons that you are short.</u>

Drawing Lewis Structures (6 & 7)

- **Step 6:** Check your structure to see if all of the atoms have their most common bonding pattern.
- **Step 7:** If necessary, try to rearrange your structure to give each atom its most common bonding pattern. One way to do this is to return to Step 2 and try another skeleton. (This step is unnecessary if all of the atoms in your structure have their most common bonding pattern.)

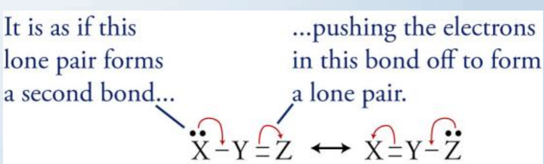
Lewis Structure Drawing Summary



Resonance

- We can view certain molecules and polyatomic ions as if they were able to resonate—to switch back and forth—between two or more different structures. Each of these structures is called a ***resonance structure***. The hypothetical switching from one resonance structure to another is called ***resonance***.

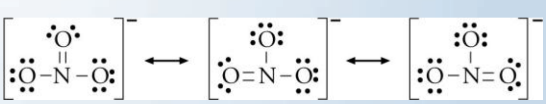
To draw Resonance Forms

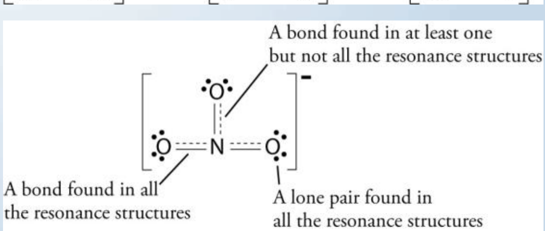


Resonance Hybrid

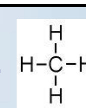
- To blend the resonance structures into a single resonance hybrid:
  - Step 1: Draw the skeletal structure, using solid lines for the bonds that are found in all of the resonance structures.
  - Step 2: Where there is sometimes a bond and sometimes not, draw a dotted line.
  - Step 3: Draw only those lone pairs that are found on every one of the resonance structures. (Leave off the lone pairs that are on one or more resonance structure but not on all of them.)

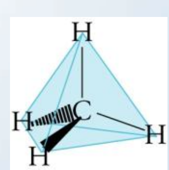
Nitrate Resonance

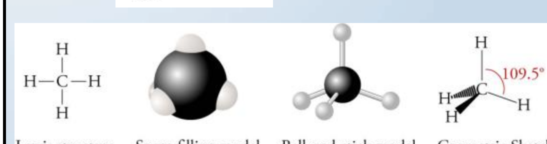




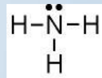
Methane, CH_4

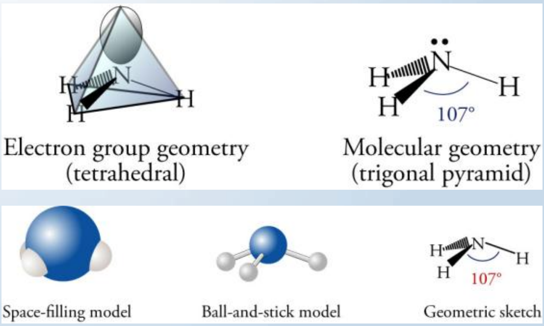




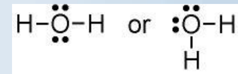


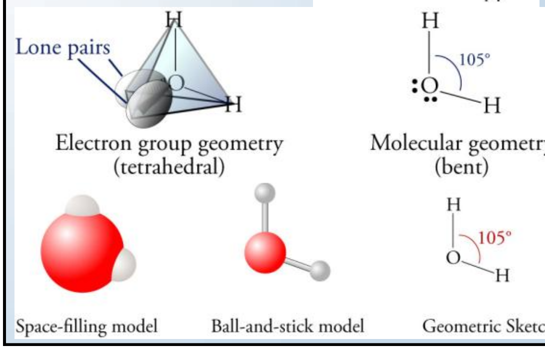
Ammonia, NH_3



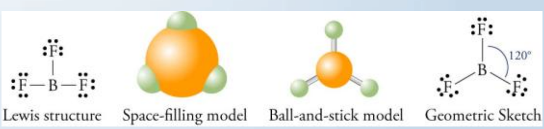


Water, H_2O

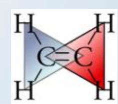


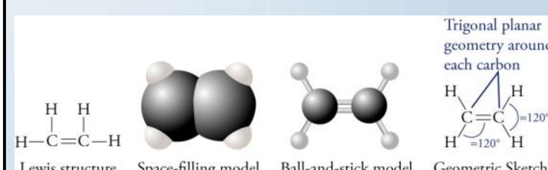


Trigonal Planar Geometry – BF_3

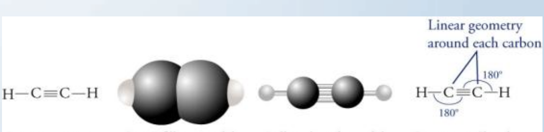


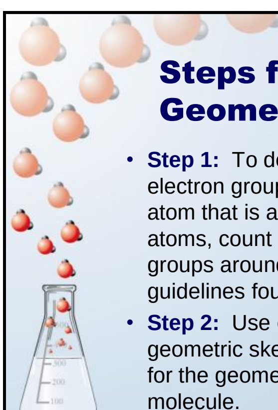
Ethene (ethylene)





Ethyne (acetylene) , C_2H_2





Steps for Molecular Geometry

- **Step 1:** To determine the name of the electron group geometry around each atom that is attached to two or more atoms, count the number of electron groups around each atom and apply the guidelines found on Table 5.2.
- **Step 2:** Use one or more of the geometric sketches shown on Table 5.2 for the geometric sketch of your molecule.

Steps for Molecular Geometry (cont.)

- **Step 3:** To determine the name of the molecular geometry around each atom that has two or more atoms attached to it, count the number of bond groups and lone pairs, and then apply the guidelines found on Table 5.2.