

NOTES - BONDING

General Rule of Thumb:

metal + nonmetal = ionic

polyatomic ion + metal or polyatomic ion = ionic (both)

nonmetal + nonmetal(s) = covalent

Intro to Ionic Bonds!

Knowing electron configurations is important because the number of valence electrons determines the chemical properties of an element.

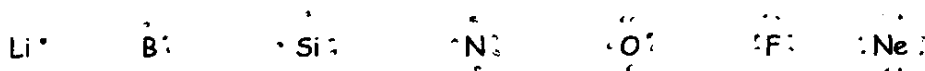
Valence Electrons: The e^- in the highest occupied energy level of an element's atoms

All elements in a particular group or family have the same number of valence electrons (and this number is equal to the group number of that element)

Examples:

Group 1 elements (Na, K, Li, H): 1 valence electron
 Group 2 elements (Mg, Ca, Be): 2 valence electrons
 Group 17 elements (Cl, F, Br): 7 valence electrons

Electron dot structures show the valence electrons as dots around the element's symbol:



Noble gas atoms are very stable; they have stable electron configurations. In forming compounds, atoms make adjustments to achieve the lowest possible (or most stable) energy.

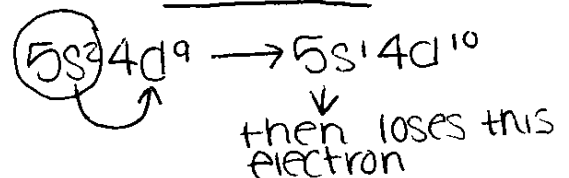
Octet rule: atoms react by changing the number of electrons so as to acquire the stable electron structure of a noble gas

- Atoms of **METALS** obey this rule by losing electrons.
 Na: 1 valence e^- \cdot
 Na: 8 valence e^- \cdot
- Atoms of **NONMETALS** obey this rule by gaining electrons.
 Cl: 7 valence e^- \cdot
 Cl: 8 valence e^- \cdot

- Transition metals are exceptions to this rule.

Example: silver (Ag) $[Kr]5s^24d^9$

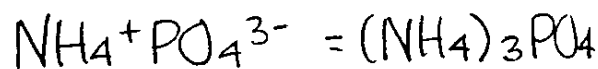
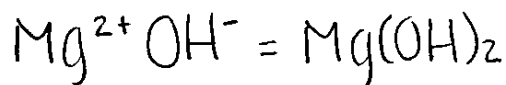
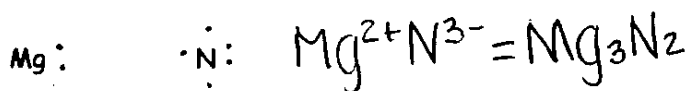
By losing one electron, it acquires a relatively stable configuration with its 4d sublevel filled (pseudo noble-gas)



IONIC BONDS / COMPOUNDS

- Anions and cations have opposite charges; they attract one another by electrostatic forces (IONIC BONDS)
- Ionic compounds are electrically neutral groups of ions joined together by electrostatic forces. (also known as salts)
- ** the positive charges of the cations must equal the negative charges of the anions.
- ** use electron dot structures to predict the ratios in which different cations and anions will combine.

Examples:

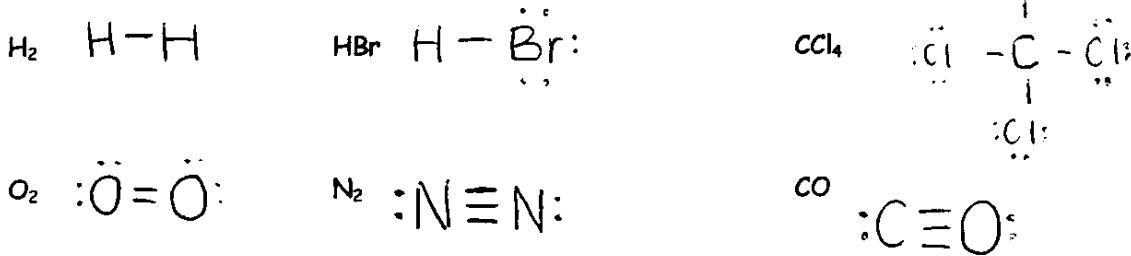


ANY TIME YOU ADD A SUBSCRIPT TO A POLYATOMIC ION, YOU MUST PUT THE ION IN PARENTHESES!

Notes: Covalent Bonds

- **Covalent bonds:** occur between 2 nonmetals; electrons are shared not transferred, as in ionic bonds)
- The result of sharing electrons is that atoms attain a more stable electron configuration.
- Most covalent bonds involve:
 - 2 electrons (single covalent bond),
 - 4 electrons (double covalent bond, or
 - 6 electrons (triple covalent bond).

- **Lewis structures** (electron dot structures) show the structure of molecules. (Bonds can be shown with dots for electrons, or with dashes: 1 dash = 2 electrons)

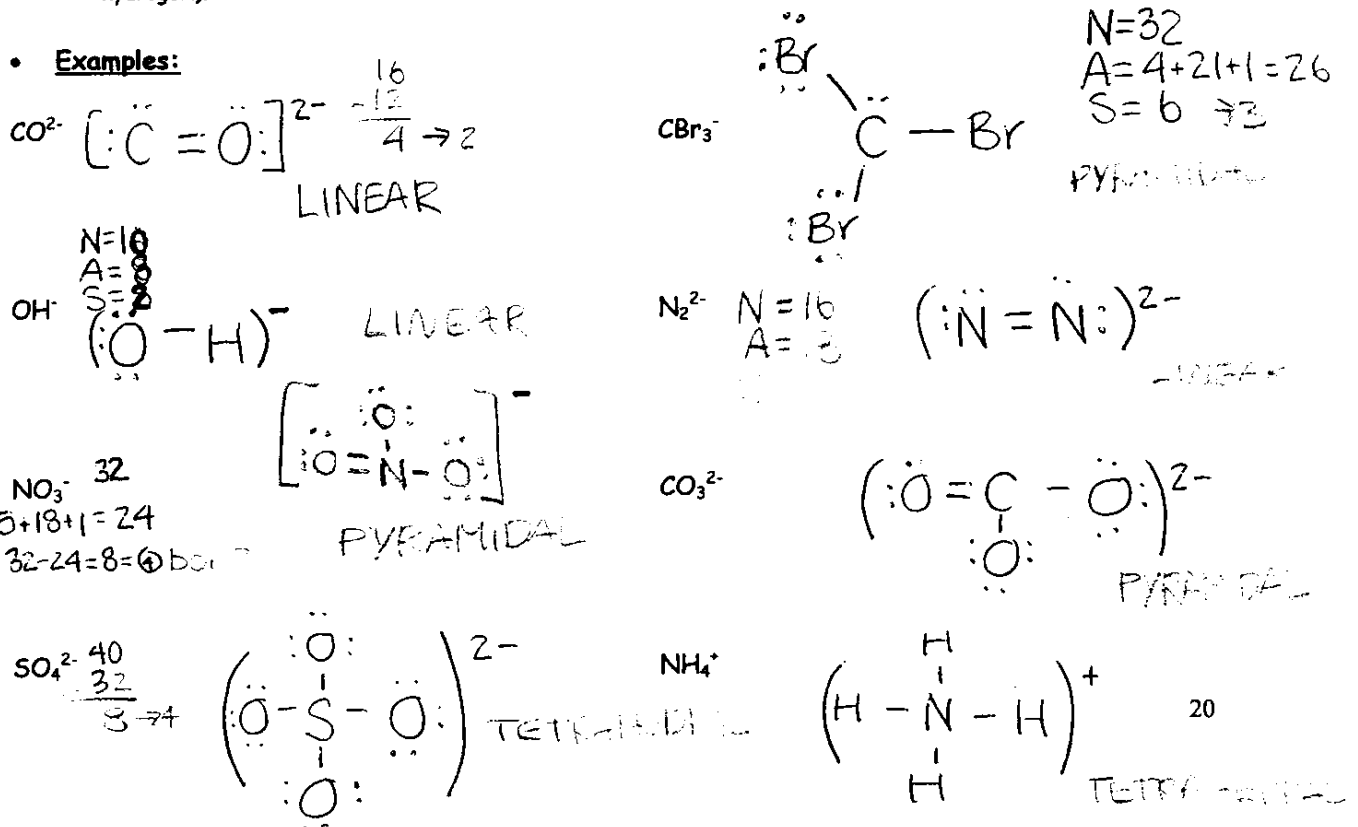


- **Octet Rule:** The representative elements achieve noble gas configurations (8 electrons) by sharing electrons.

Writing Lewis Formulas:

1. Select a skeleton for the molecule (the least electro negative element is usually the central element).
2. Calculate N (the # of valence e- need by all atoms in the molecule or polyatomic ion).
3. Calculate A (the # of electrons available).
4. Calculate S (the # of electrons shared in the molecule) $S = N - A$
5. Place the S electrons as shared pairs in the skeleton.
6. Place the additional electrons as unshared pairs to fill the octet of every representative elements (except hydrogen!).

Examples:



Notes: Electronegativity & Polarity

We've learned how valence electrons are shared to form covalent bonds between elements. So far, we have considered the electrons to be shared equally. However, in most cases, electrons are not shared equally because of a property called electronegativity.

The **ELECTRONEGATIVITY** of an element is:

The result: a "tug of war" between the nuclei of the atoms

Electronegativities are given numerical values (the most electronegative element has the highest value; the least electronegative element has the lowest value)

**See Figure 9-15 p. 263 (Glencoe)

Most electronegative element: Fluorine
(3.98)

Least electronegative element: Fr (0.70)
Cs (0.79)

*Notice the periodic trend:

-
-

**The higher the electronegativity value, the greater the ability to attract electrons to itself.

POLAR BONDS / NONPOLAR BONDS

- When the atoms in a molecule are the same, the bonding electrons are shared equally.

Result:

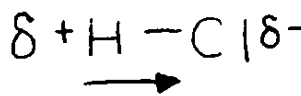
Examples:

- When 2 different atoms are joined by a covalent bond, and the bonding electrons are shared unequally, the bond is a polar covalent bond, or POLAR BOND.
- The atom with the stronger electron attraction (the more electronegative element) acquires a slightly negative charge.
- The less electronegative atom acquires a slightly positive charge.

δ = "partial"

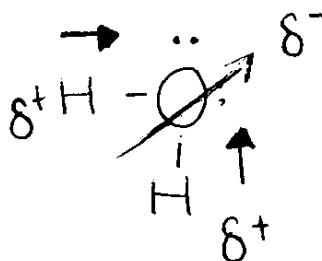
Example: HCl

Electronegativities: H = 2.20
Cl = 3.16



Example: H₂O

Electronegativities: H = 2.20
O = 3.44

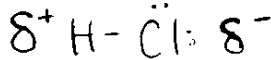


• Electronegativities help us predict the type of bond:

Electronegativity Difference	Type of Bond	Example
0.00-0.40	covalent (nonpolar)	H-H
0.41-1.00	covalent (slightly polar)	H-Cl
1.01-2.00	covalent (very polar)	H-F
2.01 or higher	ionic	Na ⁺ Cl ⁻

Polar Molecules

- A polar bond in a molecule can make the entire molecule polar
- A molecule that has 2 poles (charged regions), like H-Cl, is called a dipolar molecule, or dipole.



- The effect of polar bonds on the $\vec{\mu}$ polarity of a molecule depends on the shape of the molecule.

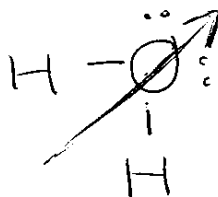
Example: CO₂



The bond polarities cancel because they are in opposite directions

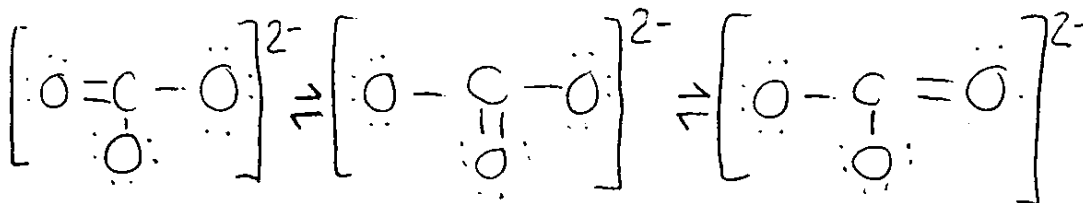
Water, H₂O, also has 2 polar bonds:

- But, the molecule is bent, so the bonds do not cancel
- H₂O is a nonpolar molecule



Notes: Resonance

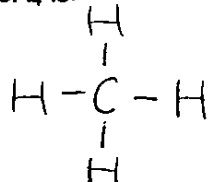
- A molecule or polyatomic ion for which 2 or more dot formulas with the same arrangement of atoms can be drawn is said exhibit RESONANCE.

Example: CO_3^{2-} 

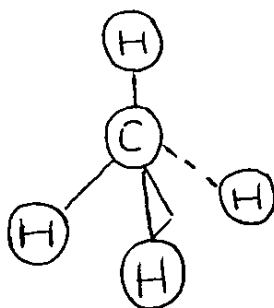
- 3 resonance structures can be drawn for CO_3^{2-}
- the relationship among them is indicated by the double arrow.
- the true structure is an average of the 3.

VSEPR Theory

- Lewis structures (electron dot structures) show the structure of molecules...but only in 2 dimensions (flat).
- BUT, molecules are 3 dimensional!
- for example, CH_4 is:



but in 3D it is:



a tetrahedron!

\triangle = coming out of page
 - - - = going into page
 — = flat on page

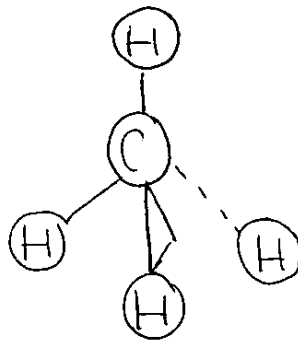
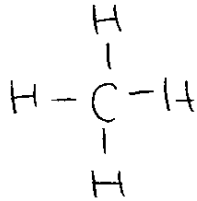
- Why do molecules take on 3D shapes instead of being flat?

Valence Shell Electron Pair Repulsion theory

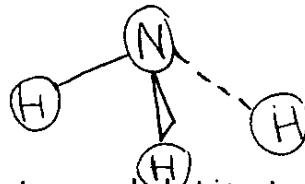
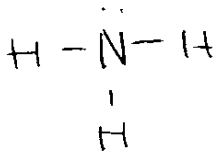
"because electron pairs repel one another, molecules adjust their shapes so that the valence electron pairs are as far apart from another as possible."

****Remember:** both shared and unshared electron pairs will repel one another.

Five basic molecule shapes:

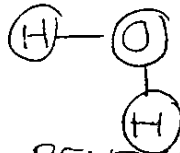
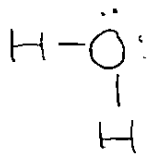
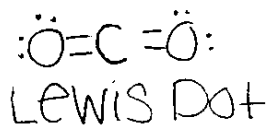
1) CH₄

TETRAHEDRAL

2) NH₃

PYRAMIDAL

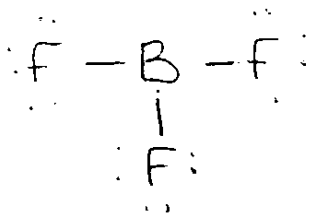
(**note: unshared pair of electron repels, but is not considered part of overall shape; no atom there to contribute to the shape)

3) H₂O4) CO₂ LEWIS DOT

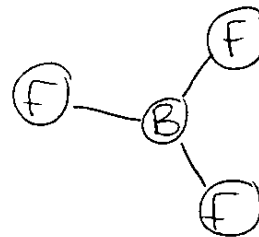
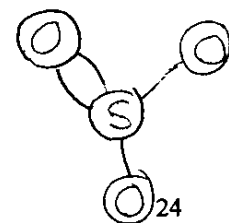
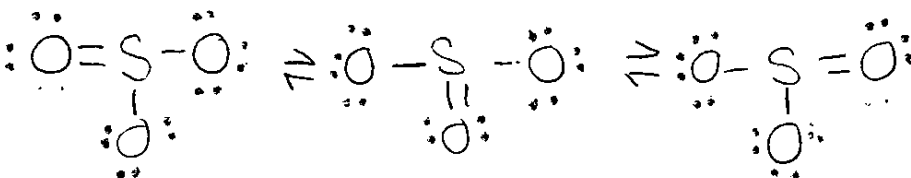
BENT/ANGULAR



LINEAR

5) BF₃

(EXCEPTION)

TRIANGULAR
PLANAR6) SO₃TRIANGULAR
PLANAR