

## Chapter 8

### *Bonding: General Concepts*

# Objectives

- ❖ To learn about ionic and covalent bonds and explain how they are formed
- ❖ To learn about the polar covalent bond
- ❖ To understand the nature of bonds and their relationship to electronegativity
- ❖ To understand bond polarity and how it is related to molecular polarity
- ❖ To learn about stable electron configurations
- ❖ To learn to predict the formulas of ionic compounds
- ❖ To learn about ionic structures
- ❖ To understand factors governing ionic size
- ❖ To learn to write Lewis structures
- ❖ To learn how to write Lewis structures for molecules with multiple bonds
- ❖ To understand molecular structure and bond angles
- ❖ To learn to predict molecular geometry from the number of electron pairs
- ❖ To learn to apply the VSEPR model to molecules with double bonds

# CHEMICAL BONDING

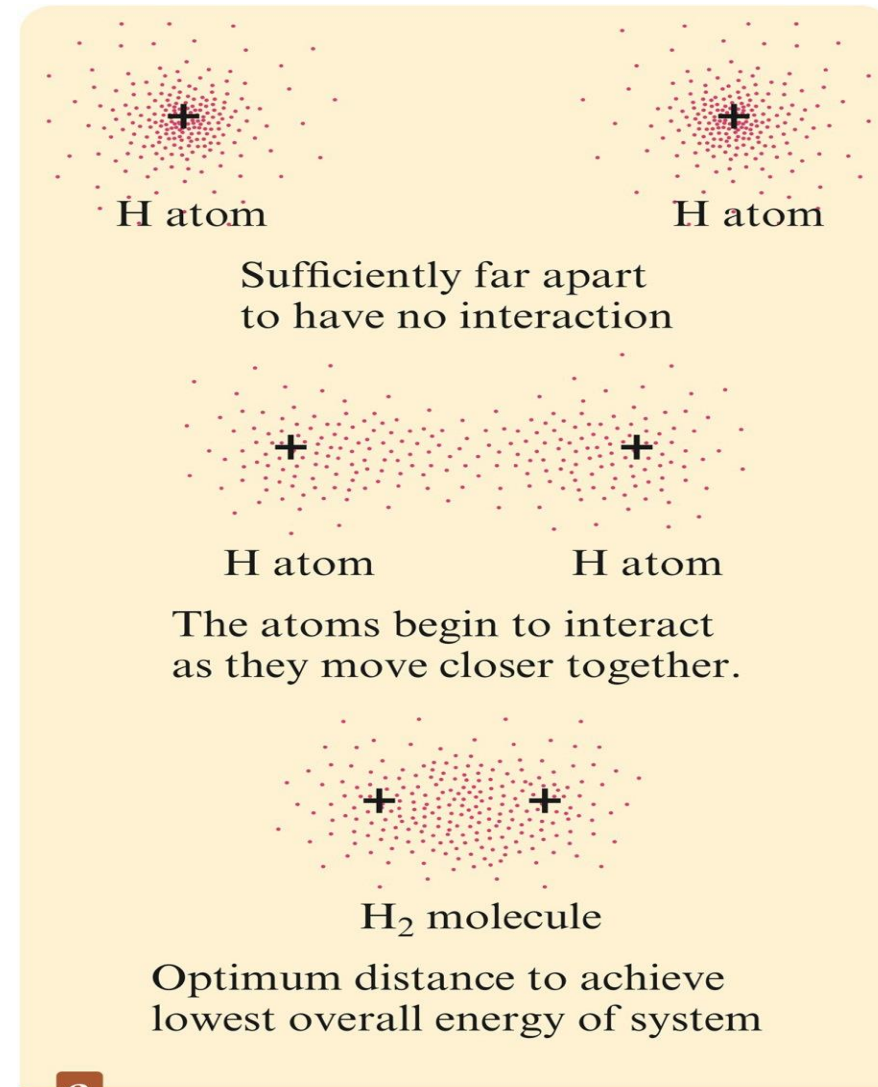
# Questions to Consider

- ❖ What is meant by the term “chemical bond”?
- ❖ Why do atoms bond with each other to form compounds?
- ❖ How do atoms bond with each other to form compounds?

# A Chemical Bond

- ❖ No simple, and yet complete, way to define this.
- ❖ Forces that hold groups of atoms together and make them function as a unit.
- ❖ A bond will form if the energy of the aggregate is lower than that of the separated atoms.
- ❖ Chemical bonding is at the heart of understanding chemistry
- ❖ Chemical bonds are strong attractive forces hold atoms and ions together in compounds/molecules
- ❖ Properties of substances are (in part) determined by chemical bonds

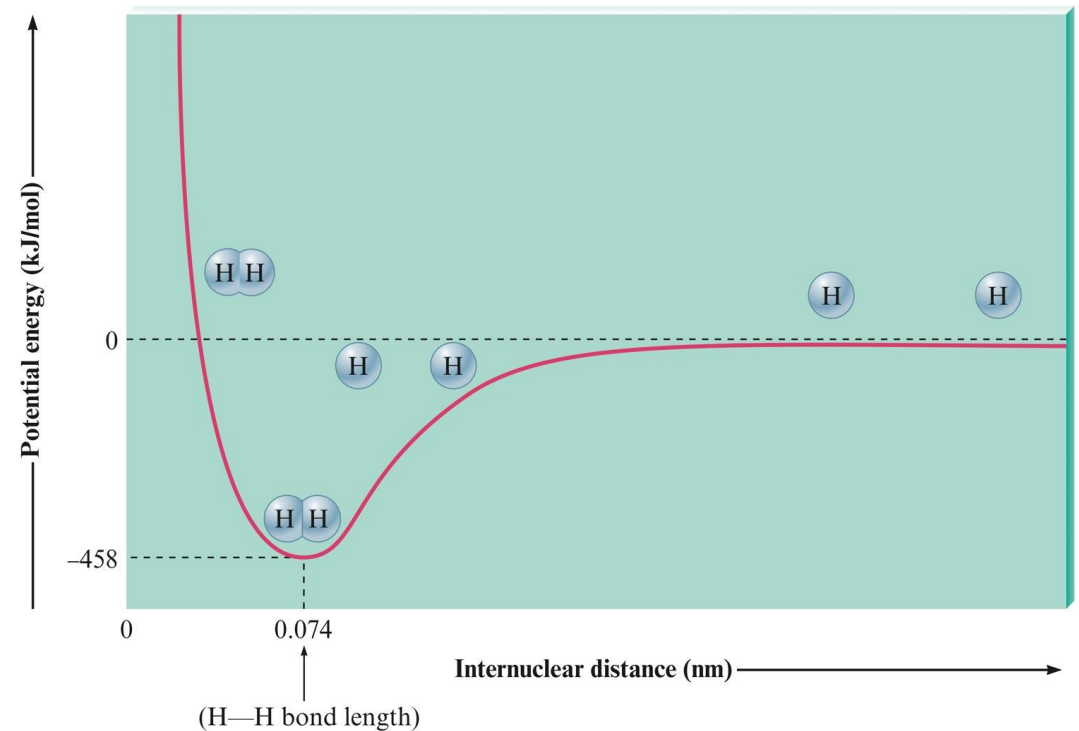
The attraction between two atoms or ions that holds them together to form chemical bond



# Fundamental questions:

## Why a chemical bond is formed?

- ❖ formed by an atom to get **more stability**. Every atom tries to get more stability by lowering its potential energy. This can be achieved by making a bond.
- ❖ The potential energy decreases when two atoms attract each other.
- ❖ Hence energy is liberated during the formation of a chemical bond i.e., it is an exothermic process.



## What is the connection between stability and electronic configuration?

- ❖ The noble gas atoms with octet configuration in their outer shells are very stable. Hence every atom tries to get octet configuration either by losing or gaining or sharing electrons. This is also called as **octet rule**.
- ❖ However this rule may not be followed always by the atoms.

# TYPES OF CHEMICAL BONDS : A BRIEF INTRODUCTION

## Key Ideas in Bonding

The chemical bonds are broadly divided into:

- ❖ **Ionic Bonding** – electrons are transferred
- ❖ **Covalent Bonding** – electrons are shared equally by nuclei
- ❖ *What about intermediate cases?*

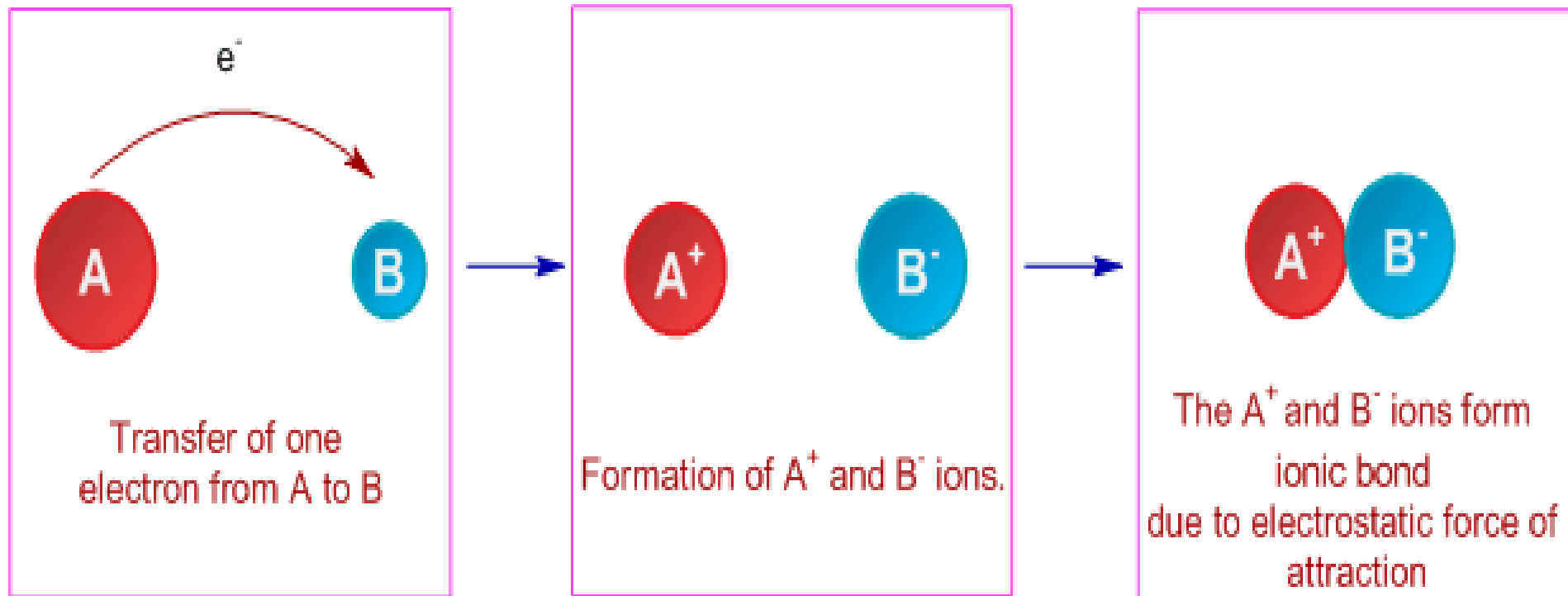
# Ionic bond

- ❖ These oppositely charged ions come closer to each other due to electrostatic force of attraction and thus by forming an ionic bond.
- ❖ An ionic bond is formed between two atoms when their electronegativity difference is **greater than 1.7 on Pauling's scale**.
- ❖ In general, an ionic bond is formed between a metal atom and a nonmetal atom. E.g., NaCl, LiF, MgCl<sub>2</sub> etc.,

## The steps involved in the formation of ionic bond can be summarized as:

- a) An electropositive atom (metal) loses electron(s) to form a positively charged ion called as cation.
- b) An electronegative atom accepts the electron(s) to form a negatively charged ion, otherwise known as anion.
- c) Thus formed oppositely charged ions come closer to each other due to electrostatic force of attraction and get stability.

The formation of ionic bond between two atoms can be visualized as follows:



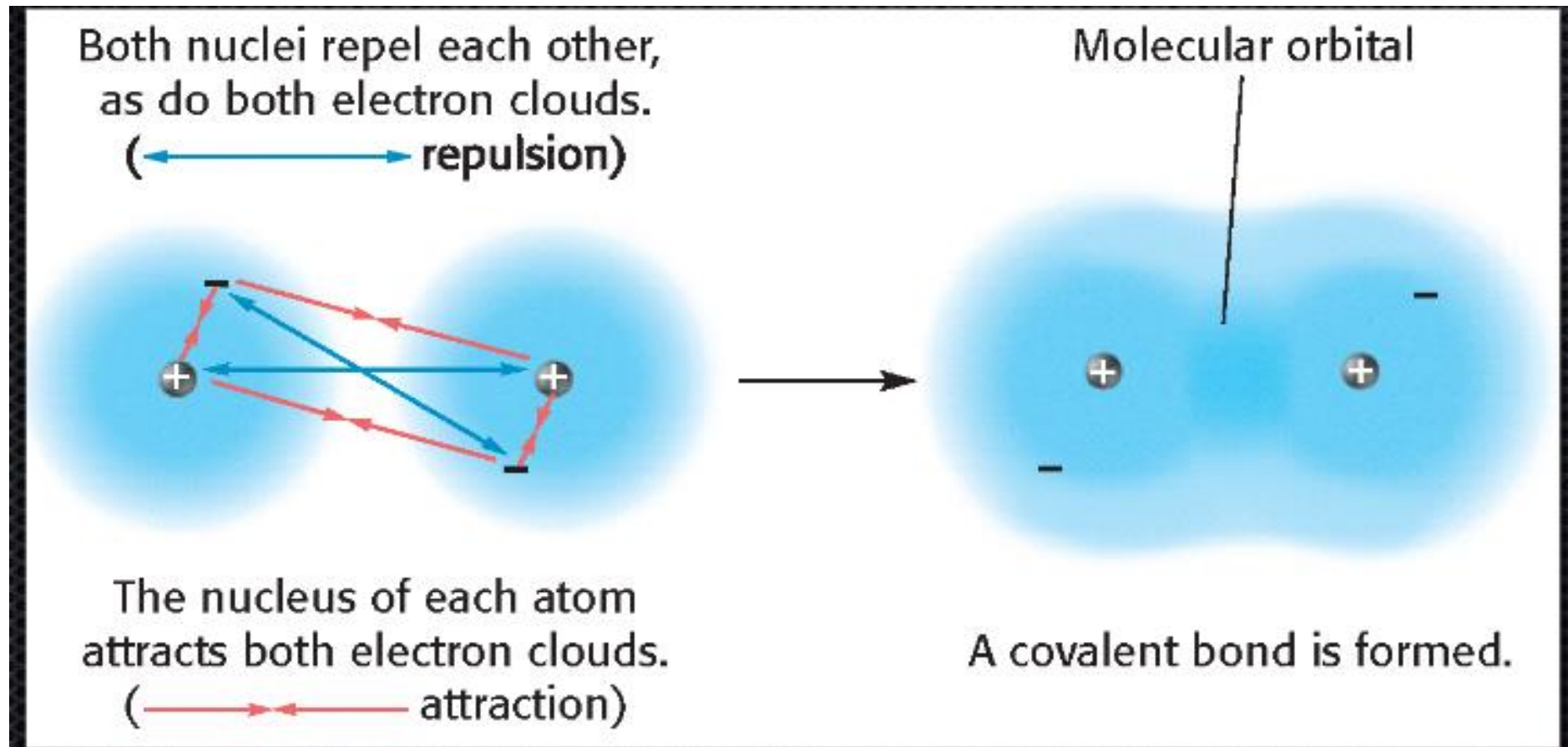
Note that there is decrease in size from A to  $A^+$  and increase in size from B to  $B^-$

# Covalent Bond

- ❖ The chemical bond formed between two atoms due to the sharing of electron pair(s) is called covalent bond.
- ❖ It is formed between two atoms for which the electronegativity difference is less than 1.7 on Pauling's scale.
- ❖ Usually two nonmetal atoms form a covalent bond. E.g.  $H_2$ ,  $F_2$ ,  $HCl$ ,  $H_2O$  etc.
- ❖ The shared pair of electrons, also known as bond pair
- ❖ It is either formed due:
  - ✓ to equal contribution of electrons by each atom participating in the bond formation; ..
  - ✓ or contribution by only one atom. This bond is also known as coordinate covalent bond or dative bond.!!

A covalent bond is formed between two atoms when their electronegativity difference is less than 1.7 on Pauling's scale.

Usually it is formed between two nonmetals.

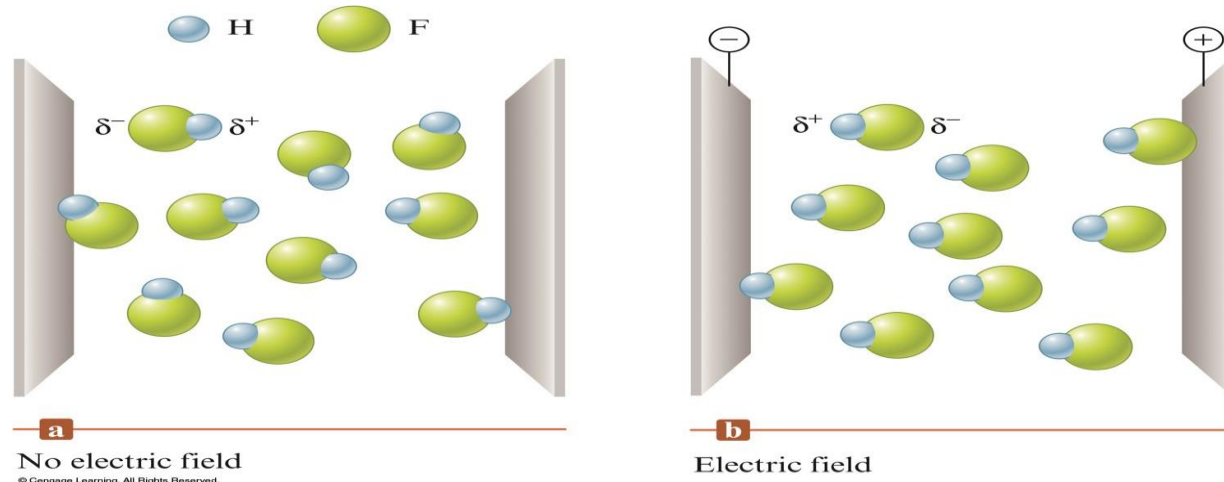


# Polar Covalent Bond

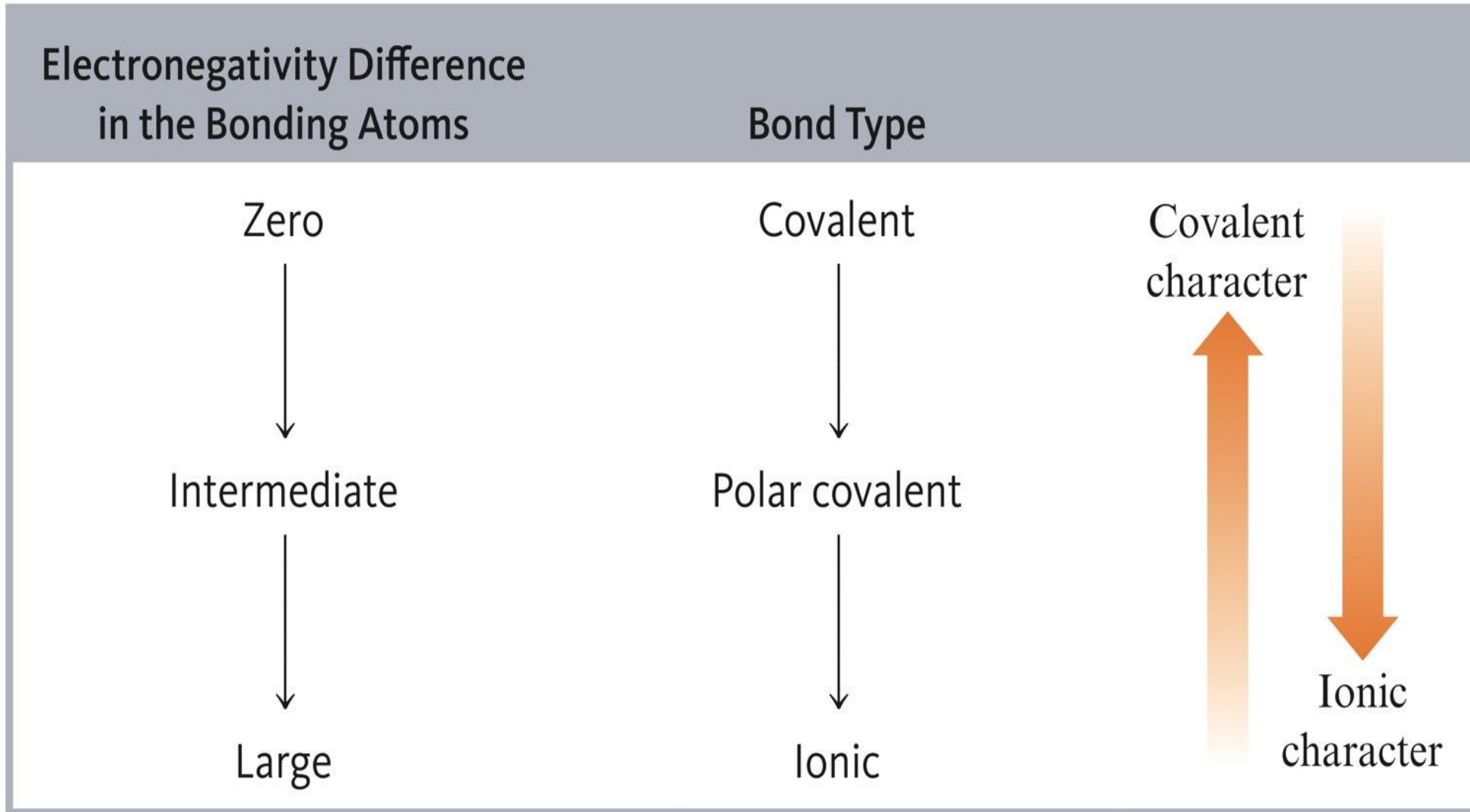
- ❖ Unequal sharing of electrons between atoms in a molecule.
- ❖ Results in a charge separation in the bond (partial positive and partial negative charge).  
 $\delta^-$  or  $\delta^+$
- ❖ The Effect of an Electric Field on Hydrogen Fluoride Molecules

For a molecule HX, the relative electronegativities of the H and X atoms are determined by comparing the measured H–X bond energy with the “expected” H–X bond energy.

**indicates a positive or negative fractional charge.**

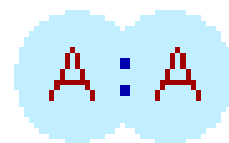


## Table 8.1 | The Relationship Between Electronegativity and Bond Type



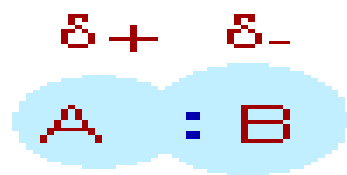
# Polarity of covalent bond

- ❖ The bond pair is *equally shared* in between two atoms *when the electronegativity difference between them is zero or nearer to zero*.
- ❖ neither of the atoms gets excess of electron density and hence carry **no** charge.
- ❖ This is called **non polar covalent bond**.
- ❖ This is true for all *homonuclear diatomic molecules*, such as  $H_2$ ,  $O_2$ ,  $N_2$ ,  $F_2$ , and  $Cl_2$ , because the two identical atoms have identical electronegativities.



non polar covalent bond

- ❖ When there is a considerable difference in the electronegativity,..
- ❖ The bond pair is no longer shared equally between the atoms.
- ❖ It is shifted slightly towards the atom with higher electronegativity by creating **partial negative charge** (represented by  $\delta^-$ ) over it.
- ❖ the atom with less electronegativity gets partial positive charge (represented by  $\delta^+$ ).
- ❖ This type of bond is also referred to as **polar covalent bond**.



polar covalent bond

Note: B is more  
electronegative

# Covalent bonds can have ionic character

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## Electronegativity

<u>Atoms</u>	<u>Difference</u>	<u>Type of Bond</u>
N-N	$3.0 - 3.0 = \mathbf{0.0}$	Nonpolar covalent
Cl-Br	$3.0 - 2.8 = \mathbf{0.2}$	Nonpolar covalent
H-Si	$2.1 - 1.8 = \mathbf{0.3}$	Nonpolar covalent

- ▶ A **polar bond** has separate centers of positive and negative charge.

# Polar Bond

- ❖ A *molecule* with separate centers of positive and negative charge is a *polar molecule*.
- ❖ The difference in electronegativity between two atoms gives a measure of the polarity between two atoms.
- ❖ A polar molecule has a polar bond with a moderate electronegativity difference (0.5 to 1.7) between the two atoms involved in the polar bond.
- ❖ In HF, in which the electron pairs are shared unequally are called polar covalent bonds can be represented as:



- ❖ This means that the F end of the molecule is somewhat more negative than the H end.

# Electric Dipole

- ❖ The separation of charge in a polar covalent bond creates an **electric dipole**.
- ❖ We expect the dipoles in the covalent molecules HF, HCl, HBr, & HI to be different because F, Cl, Br, and I have different electronegativities.
- ❖ This tells us that atoms of these elements have different **tendencies to attract an electron pair that they share with hydrogen**.

	Most polar			Least polar
	→	→	→	→
	H—F	H—Cl	H—Br	H—I
EN:	2.1 4.0	2.1 3.0	2.1 2.8	2.1 2.5
$\Delta(\text{EN})$	1.9	0.9	0.7	0.4

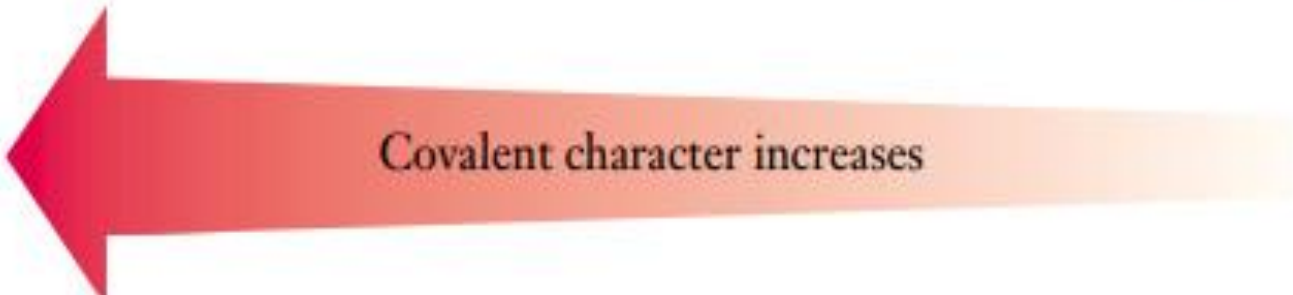
- ❖ The longest arrow indicates the largest dipole, or greatest separation of electron density in the molecule

In summary, we can describe chemical bonding as a continuum that may be represented as

$\Delta(\text{EN})$ for the bonding atoms	zero	→	intermediate	→	large
Bonding types	nonpolar covalent	→	polar covalent	→	ionic



Ionic character increases



Covalent character increases

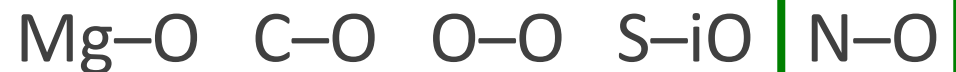
**EXAMPLE!**

Arrange the following bonds from most to least polar:

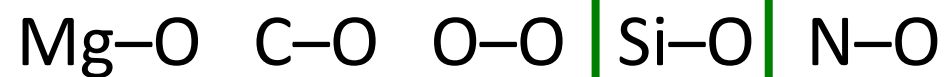


## CONCEPT CHECK!

Which of the following bonds would be the least polar yet still be considered polar covalent?



Which of the following bonds would be the most polar without being considered ionic?



# DIPOLE MOMENTS

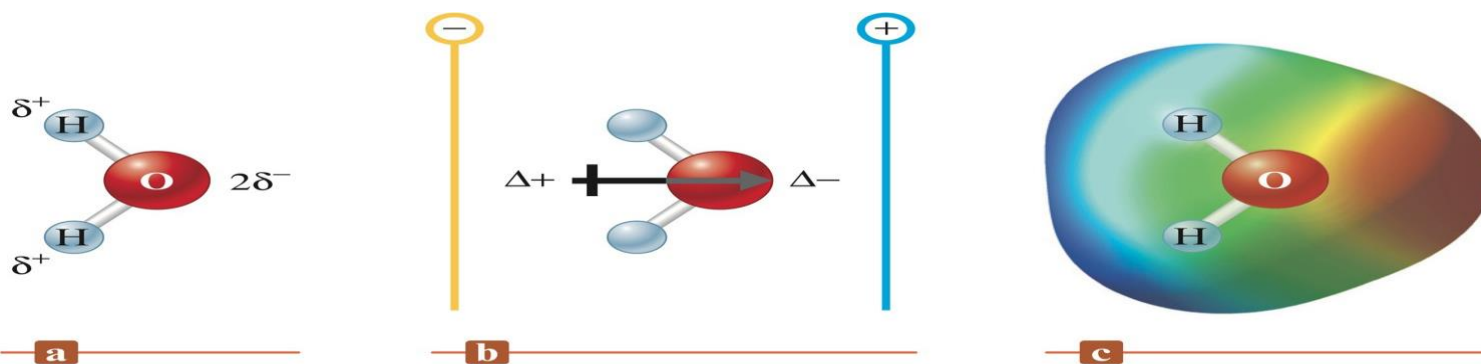
- ❖ We indicate the polarity of a molecule by its **dipole moment**, which measures the separation of charge within the molecule.
- ❖ The **dipole moment** ( $\mu$ ) of a molecule is the product of the magnitude of the charge ( $q$ ) and the distance ( $d$ ) that separates the centers of positive and negative charge. **Note:**  $q$  is Coulomb and  $d$  in m

$$\mu = d \times q$$

- ❖ A unit of dipole moment is the **debye (D)**. One debye (D) is equal to  $3.34 \times 10^{-30}$  C m.

# Bond Polarity and Dipole Moment

- ❖ Property of a molecule whose charge distribution can be represented by a center of positive charge and a center of negative charge.
- ❖ Bond dipoles have both a **magnitude** and a **direction** (they are **vector** quantities).
- ❖ Use an arrow to represent a dipole moment.
  - ✓ Point to the negative charge center with the tail of the arrow indicating the positive center of charge.

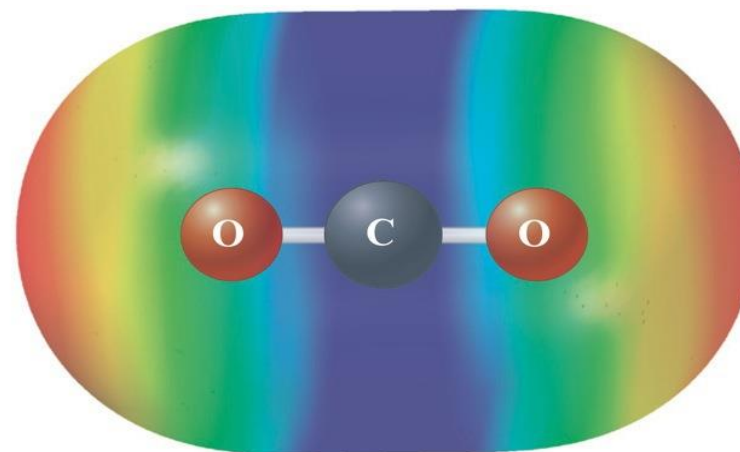


# No Net Dipole Moment (Dipoles Cancel)

- ❖ Ordinarily, a polar molecule must have polar bonds, *BUT ... polar bonds are not sufficient.*
- ❖ *A molecule may have polar bonds and be a nonpolar molecule – IF the bond dipoles cancel*

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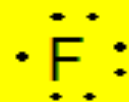
**b****c**

# Localized Electron Bonding Model

- ❖ A molecule is composed of atoms that are bound together by sharing pairs of electrons using the atomic orbitals of the bound atoms
  - ❖ Electron pairs are assumed to be localized on a particular atom or in the space between two atoms:
    - *Lone pairs* – pairs of electrons localized on an atom
    - *Bonding pairs* – pairs of electrons found in the space between the atoms
1. Description of valence electron arrangement (Lewis structure).
  2. Prediction of geometry (VSEPR model).
  3. Description of atomic orbital types used by atoms to share electrons or hold lone pairs.

# LEWIS STRUCTURE

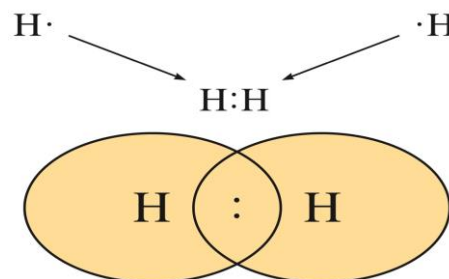
- ❖ **Lewis Theory** was Formulated by Gilbert Newton Lewis and other two scientists (Walther Kossel, and Irving Langmuir)
- ❖ They came up with a theory to explain chemical bonding (**Between 1916 and 1919**)
- ❖ Lewis structures are visual representations of the bonds between atoms and illustrate the lone pairs of electrons in molecules
- ❖ A **Lewis structure** (or Lewis dot symbols) is a chemical symbol of an element surrounded by dots, each representing one of the s and/or p valence electrons of the atom.
- ❖ The electrons in the valence shell of the atom are shown as dots around it



- ❖ The **Lewis Structure** Shows how valence electrons are arranged among atoms in a molecule.
- ❖ Reflects central idea that stability of a compound relates to noble gas electron configuration.

## Duet Rule

- ✓ Hydrogen forms stable molecules where it shares two electrons



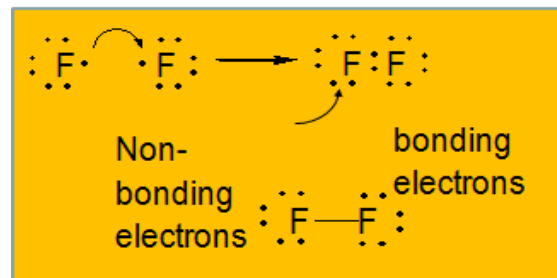
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## Octet Rule

- ✓ Elements form stable molecules when surrounded by eight electrons



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## Single Covalent Bond

- ❖ A covalent bond in which two atoms share one pair of electrons.



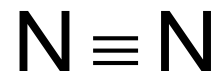
## Double Covalent Bond

- ❖ A covalent bond in which two atoms share two pairs of electrons.



## Triple Covalent Bond

- ❖ A covalent bond in which two atoms share three pairs of electrons.



# Steps for Writing Lewis Structures

1. Sum the valence electrons from all the atoms.
2. Identify the central atom, its often the atom with the lowest electronegativity. Hydrogen, is never a central atom.
  - ✓ The least electronegative atom is usually the one requiring the most electrons to fill its octet.
3. Write the skeletal structure, & Use a pair of electrons to form a bond in the structure by single covalent bonds
4. For each single bond thus formed, subtract two from the total number of the valence electrons.
5. Atoms usually have noble gas configurations. Arrange the remaining electrons to satisfy the octet rule (or duet rule for hydrogen).
6. If at this stage, the central atom (s) lacks an octet, form multiple covalent bonds by converting lone pair electrons from terminal atoms into bond pairs.
7. Oxygen atoms do not bond to each other except in  $O_2$  and  $O_3$  molecules; hydrogen peroxide,  $H_2O_2$ , and its derivatives, the peroxides, which contain the  $O_2^{2-}$  group

# Steps for Writing Lewis Structures

Example:  $\text{H}_2\text{O}$

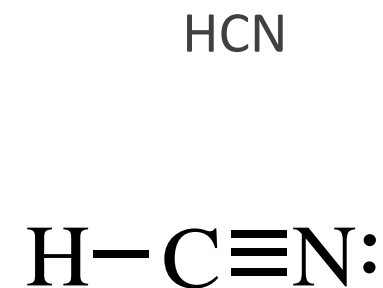
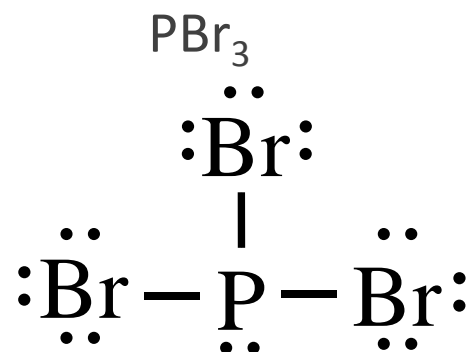
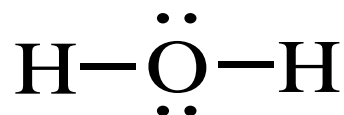
1. Sum the valence electrons from all the atoms. (Use the periodic table.)

$$2 (1 e^-) + 6 e^- = 8 e^- \text{ total}$$

2. Use a pair of electrons to form a bond between each pair of bound atoms.



3. Atoms usually have noble gas configurations. Arrange the remaining electrons to satisfy the octet rule (or duet rule for hydrogen).



# Steps for Writing Lewis Structures

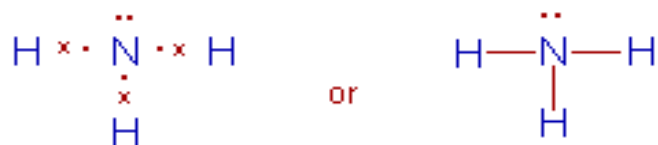
## Example: Methane (CH<sub>4</sub>):

- ✓ The C atom forms 4 covalent bonds by contributing four of its valence electrons. It forms 4 bonds with four H atoms. Thus it gets octet configuration



## Example: Ammonia (NH<sub>3</sub>)

- ✓ the nitrogen atom contributes 3 of its valence electrons to form three bond pairs which are shared with hydrogen atoms.



# Steps for Writing Lewis Structures

## For Multiple Covalent Bonds

### Example: Dinitrogen (N<sub>2</sub>)

- ✓ each nitrogen atom contributes 3 electrons to form 3 bond pairs, which in turn are shared by two nitrogen atoms. Each nitrogen also contains one lone pair.



### Example: Carbon dioxide(CO<sub>2</sub>):

- ✓ The carbon atom contributes four of its valence electrons, whereas each oxygen atom contributes two electrons
- ✓ There are two electron pairs shared between carbon and one of the oxygen atom i.e., a double bond, C=O is formed. There are two such C=O bonds in CO<sub>2</sub> molecule



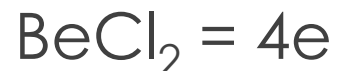
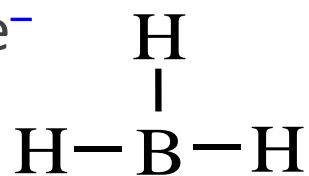
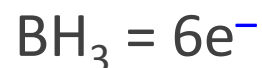
**CONCEPT CHECK!**

Draw a Lewis structure for each of the following molecules:

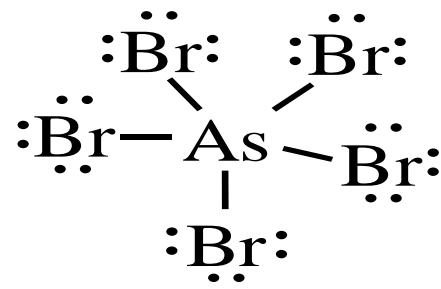
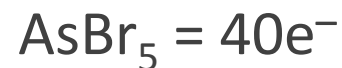
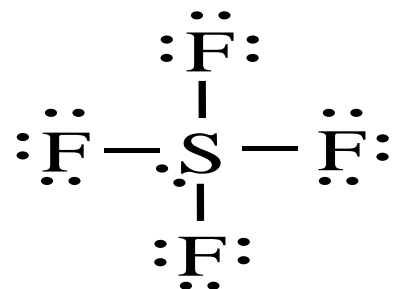
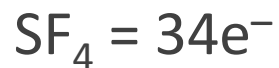


# Exceptions of the Octet Rule

- Boron and Beryllium tends to form compounds in which the atoms has fewer than eight electrons around it (it does not have a complete octet).



- When it is necessary to exceed the octet rule for one of several third-row (or higher) elements, place the extra electrons on the central atom.



**CONCEPT CHECK!**

Draw a Lewis structure for each of the following molecules:

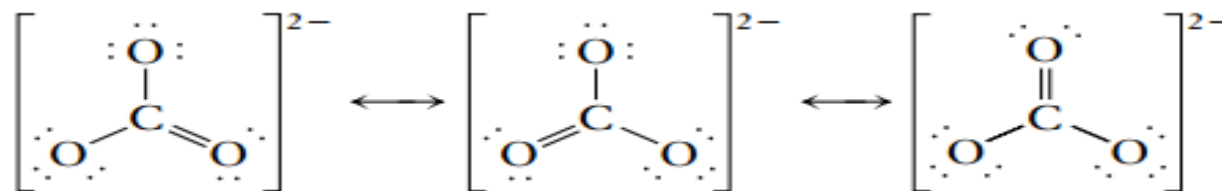
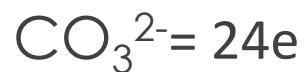
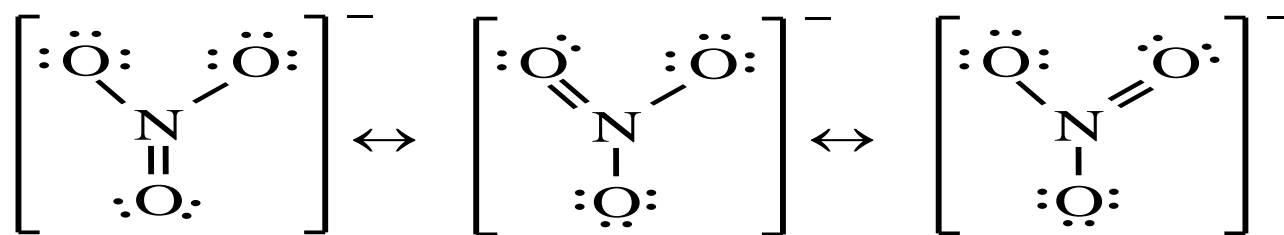
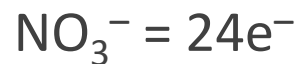


# Let's Review

- ❖ C, N, O, and F should always be assumed to obey the octet rule.
- ❖ B and Be often have fewer than 8 electrons around them in their compounds.
- ❖ Second-row elements never exceed the octet rule.
- ❖ Third-row and heavier elements often satisfy the octet rule but can exceed the octet rule by using their empty valence *d* orbitals.
- ❖ When writing the Lewis structure for a molecule, satisfy the octet rule for the atoms first. If electrons remain after the octet rule has been satisfied, then place them on the elements having available *d* orbitals (elements in Period 3 or beyond).

# Resonance

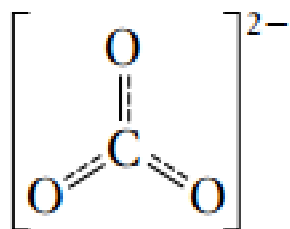
- ❖ More than one valid Lewis structure can be written for a particular molecule.



- ❖ Actual structure is an average of the resonance structures.
- ❖ Electrons are really delocalized – they can move around the entire molecule.

# Resonance

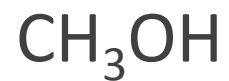
- ❖ The three structures above are resonance structures of the carbonate ion.
- ❖ The relationship among them is indicated by the double-headed arrows.
- ❖ This symbol does not mean that the ion flips back and forth among these three structures. The true structure can be described as an average, **or hybrid**, The true structure of the three.



(lone pairs on O atoms not shown)

**CONCEPT CHECK!**

Draw a Lewis structure for each of the following molecules:



# Formal Charge

- ❖ Used to evaluate nonequivalent Lewis structures.
- ❖ Atoms in molecules try to achieve formal charges as close to zero as possible.
- ❖ Any negative formal charges are expected to reside on the most electronegative atoms & positive on the least electronegative atoms
  - ❖ Formal charge = (# valence e<sup>-</sup> on free neutral atom) – (# valence e<sup>-</sup> assigned to the atom in the molecule).

## Calculation

$$\text{FC} = \text{No. valence electrons in isolated atom} - \text{No. bonds to atom} - \text{No. unshared electrons}$$

- ❖ Assume:
  - Lone pair electrons belong entirely to the atom in question.
  - Shared electrons are divided equally between the two sharing atoms.

# Rules Governing Formal Charge

To calculate the formal charge on an atom:

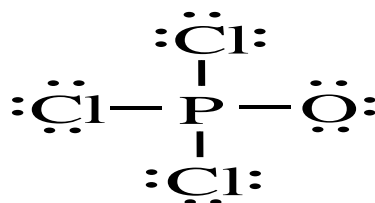
1. Take the sum of the lone pair electrons and one-half the shared electrons.
2. Subtract the number of assigned electrons from the number of valence electrons on the free, neutral atom.
3. The sum of the formal charges of the atoms in the Lewis structure must equal to zero for a neutral atom or equal to the ionic charge for a polyatomic ion

**Example:** Consider the Lewis structure for  $\text{POCl}_3$ . Assign the formal charge for each atom in the molecule.

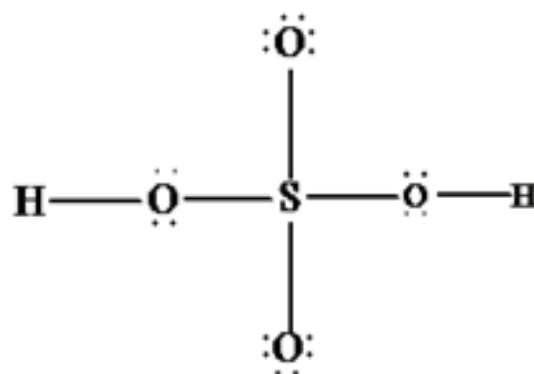
$$\text{P: } 5 - 4 = +1$$

$$\text{O: } 6 - 7 = -1$$

$$\text{Cl: } 7 - 7 = 0$$



Find the formal charges of all the atoms in the sulphuric acid structure.



$$\text{Formal charge on S} = 6 - (4 + 0) = +2$$

An isolated S atom has 6 electrons

There are no. unshared pairs on the S

$$\text{Sulfur: FC} = 6 - (4 + 0) = +2$$

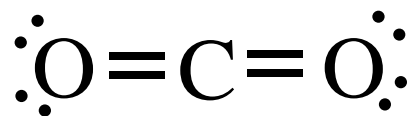
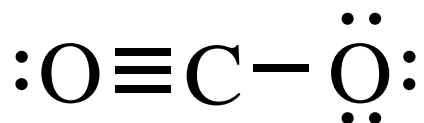
$$\text{Hydrogen: FC} = 1 - (1 + 0) = 0$$

$$\text{Hydroxyl oxygen: FC} = 6 - (2 + 4) = 0$$

$$\text{Oxygen not bonded to hydrogen: FC} = 6 - (1 + 6) = -1$$

# Rules Governing Formal Charge

- ❖ The sum of the formal charges of all atoms in a given molecule or ion must equal the overall charge on that species.
- ❖ If nonequivalent Lewis structures exist for a species, those with formal charges closest to zero and with any negative formal charges on the most electronegative atoms are considered to best describe the bonding in the molecule or ion.



# DRAWBACKS OF LEWIS THEORY

- ❖ Lewis theory could not explain the geometry of molecules & bond angle in them.
- ❖ It could not explain why some molecules are violating the octet rule.
- ❖ This is a qualitative explanation for **covalent bond only**.
- ❖ To fulfill these gaps and to explain the covalent bond formation quantitatively, the **Valence bond theory (VBT)** was put forward.
- ❖ However it is important to learn **VSEPR theory**, another qualitative model, which was put forward to explain the shapes of molecules, before moving on to Valence bond theory.

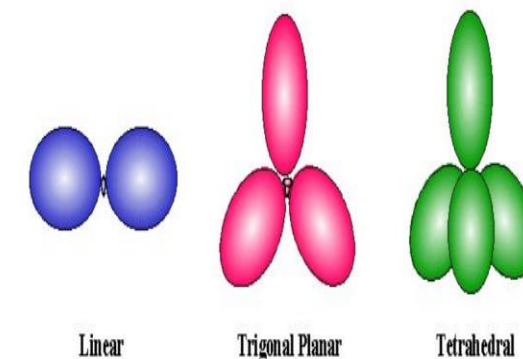
# Molecular Structure: The VSEPR Model

## VALENCE SHELL ELECTRON PAIR REPULSION (VSEPR) THEORY

- ▶ Sidgwick and Powell in 1940, proposed a simple theory based on the repulsive interactions of the electron pairs in the valence shell of the atoms
- ▶ Helps us to understand & predict the spatial arrangement of atoms in a polyatomic molecule or ion
- ▶ VSEPR theory uses Lewis structures to predict the molecular geometry of covalently bonded molecules (molecular shape)
- ▶ The structure around a given atom is determined principally by minimizing electron pair repulsions in the valence shell.
- ▶ Molecular shape offers insight into a wide range of important physical properties (polarity, solubility, volatility, chirality, etc).

Balloons tied together adopt arrangements which minimize steric clashes between neighbors:

Low energy arrangements of balloons!



## The main postulates of VSEPR theory are as follows:

- ❖ The electron pairs in the valence shell around the **central atom** of a molecule **repel each other** & tend to orient in space so as to minimize the repulsions & maximize the distance between them.
- ❖ There are two types of valence shell electron pairs viz.
  - ✓ **Bond pairs** are shared by two atoms & are attracted by two nuclei. Hence they occupy less space and cause less repulsion.
  - ✓ **Lone pairs** are not involved in bond formation & are in attraction with only one nucleus. Hence they occupy more space. As a result, the lone pairs cause more repulsion.
- ❖ The order of repulsion between different types of electron pairs is as follows:  
**Lone pair - Lone pair > Lone Pair - Bond pair > Bond pair - Bond pair**

**Note:** The bond pairs are usually represented by a solid line, whereas the lone pairs are represented by a lobe with two electrons.

## The main postulates of VSEPR theory are as follows:

- ❖ In VSEPR theory, the multiple bonds are treated as if they were single bonds.
  - ✓ The electron pairs in multiple bonds are treated collectively as a **single super pair**.
  - ✓ The repulsion caused by bonds increases with increase in the # of bonded pairs between two atoms i.e.,
  - ✓ A triple bond causes more repulsion than a double bond which in turn causes more repulsion than a single bond.
- ❖ The shape of a molecule can be predicted from the number and type of valence shell electron pairs around the central atom.
  - ✓ When the valence shell of central atom contains only bond pairs, the molecule assumes symmetrical geometry due to even repulsions between them.
  - ✓ However the symmetry is distorted when there are also lone pairs along with bond pairs due to uneven repulsion forces.
  - ✓ The best spatial arrangement of the bonding pairs of electrons in the valence orbitals is one in which the repulsions are minimized

## Steps to Apply the VSEPR Model

1. Draw the Lewis structure for the molecule.
2. Count total number of electron pairs around the central atom. Arrange them to minimize the electron shell repulsion
3. Count the electron pairs and arrange them in the way that minimizes repulsion (put the pairs as far apart as possible).
4. Determine the positions of the atoms from the way electron pairs are shared (how electrons are shared between the central atom and surrounding atoms).
5. Determine the name of the molecular structure from positions of the atoms.

**Example!**

Determine the shape for each of the following molecules, and include **bond angles**:



✓ **Solutions**

HCN – linear, 180°

PH<sub>3</sub> – trigonal pyramid, 109.5° (107°)

SF<sub>4</sub> – see saw, 90°, 120°



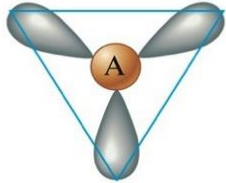
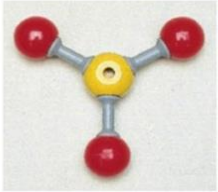
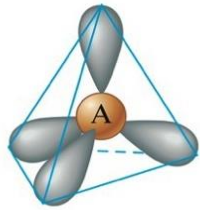
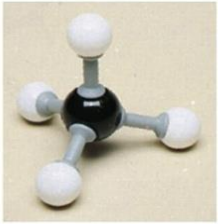
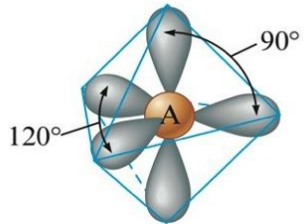
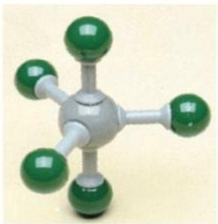
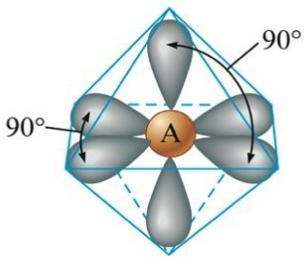
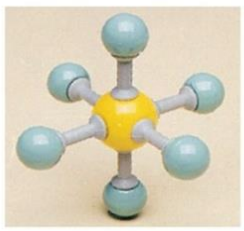
O<sub>3</sub> – bent, 120°

KrF<sub>4</sub> – square planar, 90°, 180°

## *Let's Think About It*

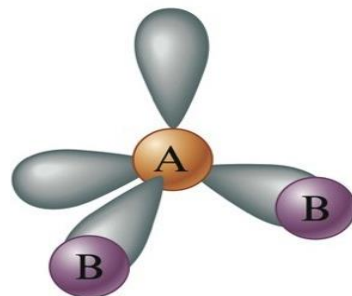
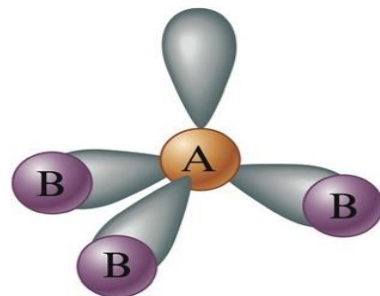
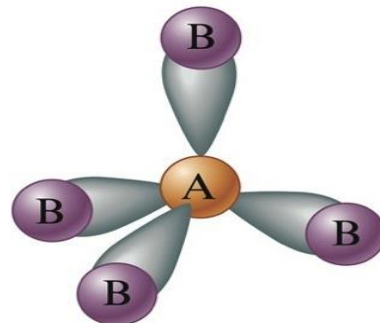
- ❖ Draw the Lewis structure for  $\text{XeF}_4$ .
- ❖ Does  $\text{XeF}_4$  contain lone pairs?
- ❖ Is the molecule polar or nonpolar overall? Why?

# Arrangements of Electron Pairs Around an Atom Yielding Minimum Repulsion

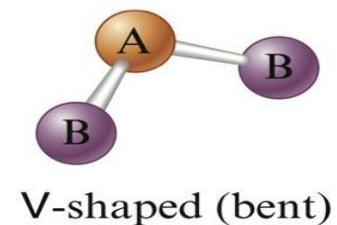
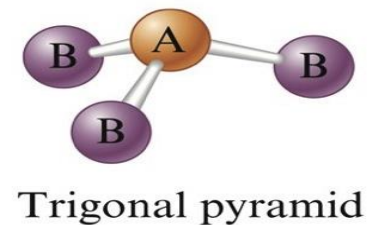
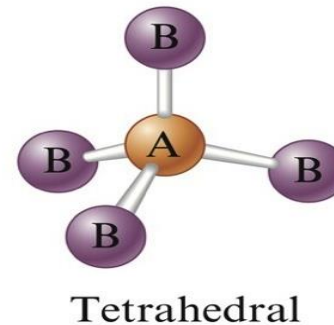
Number of Electron Pairs	Arrangement of Electron Pairs	Example
2	Linear 	
3	Trigonal planar 	
4	Tetrahedral 	
5	Trigonal bipyramidal 	
6	Octahedral 	

# Structures of Molecules That Have Four Electron Pairs Around the Central Atom

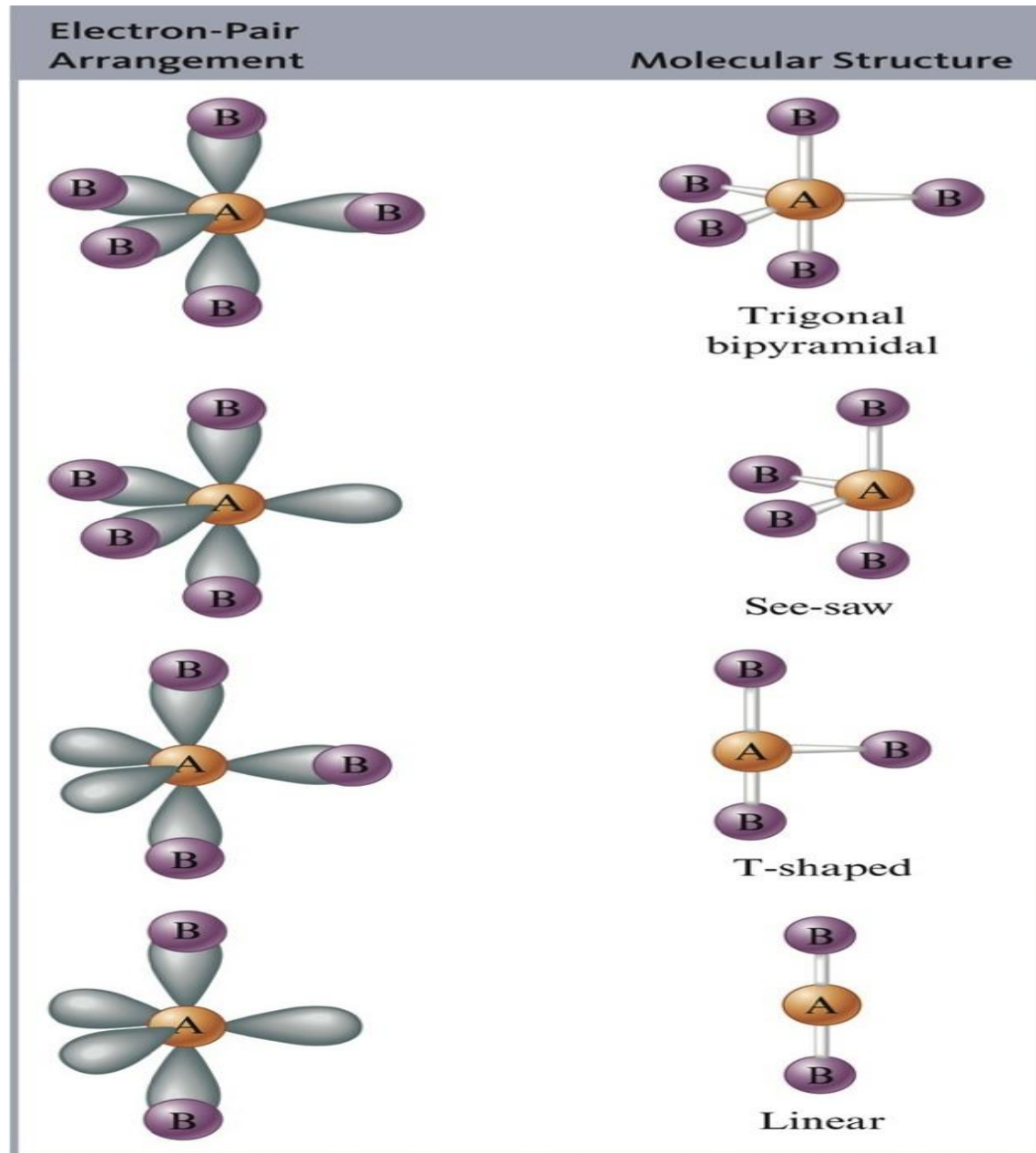
Electron-Pair Arrangement



Molecular Structure



# Structures of Molecules with Five Electron Pairs Around the Central Atom



## RELATION BETWEEN NUMBER & TYPE OF VALENCE ELECTRON PAIRS WITH THE SHAPE OF MOLECULE

- ❖ The shape of molecule and also the approximate bond angles can be predicted from the number and type of electron pairs in the valence shell of central atom as tabulated below.

In the following table the molecule is represented by "AXE" notation, where:


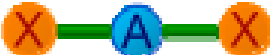
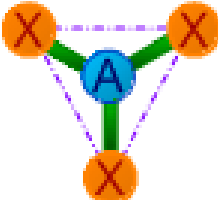
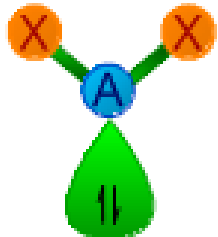
A = Central atom

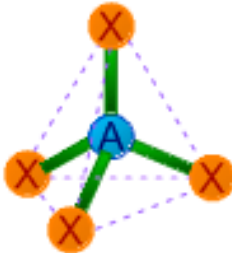
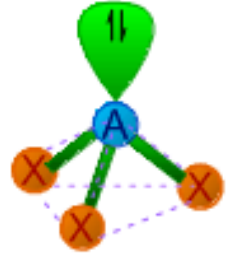
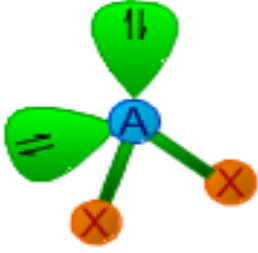
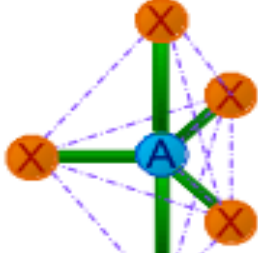
X = atom bonded to the central atom either by a single bond or by multiple bond; **indicating a bond pair.**

E = Lone pair

### Note:

- ❖ The sum of # of ligand atoms (**X**) and # of lone pairs (**E**) is also known as **steric number**.
  - ✓ The bond pairs are shown as green colored thick lines,
  - ✓ whereas the lone pairs are shown as point charges using green colored lobes.

<u>Steric number</u>	Number of Bond pairs	Number of Lone pairs	Formula	Shape of molecule	Approximate Bond angles	Examples	
1	1	0	AX	Linear		-	ClF, BrF, BrCl, HF, O <sub>2</sub>
2	2	0	AX <sub>2</sub>	Linear		180°	BeCl <sub>2</sub> , HgCl <sub>2</sub> , CO <sub>2</sub>
3	3	0	AX <sub>3</sub>	Trigonal planar		120°	BF <sub>3</sub> , CO <sub>3</sub> <sup>2-</sup> , NO <sub>3</sub> <sup>-</sup> , SO <sub>3</sub>
	2	1	AX <sub>2</sub> E	Angular		120°	SO <sub>2</sub> , SnCl <sub>2</sub> , O <sub>3</sub> , NSF, NO <sub>2</sub> <sup>-</sup>

4	0	$AX_4$	Tetrahedral		$109^\circ 28'$	$CH_4, SiCl_4,$ $NH_4^+,$ $PO_4^{3-},$ $SO_4^{2-},$ $ClO_4^-$
3	1	$AX_3E$	<u>Trigonal</u> pyramidal		around $109^\circ 28'$	$NH_3, PCl_3,$ $XeO_3$
2	2	$AX_2E_2$	Angular		around $109^\circ 28'$	$H_2O, SCl_2,$ $Cl_2O,$  $OF_2$
5	0	$AX_5$	<u>Trigonal</u> bipyramidal		$120^\circ$ & $90^\circ$	$PCl_5, SOF_4$

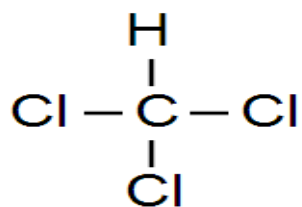
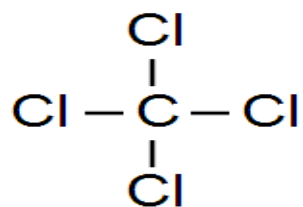
# Molecular Shapes, Dipole Moments and polarity of molecules

## To predict molecular polarity:

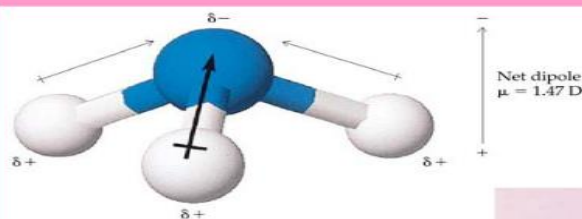
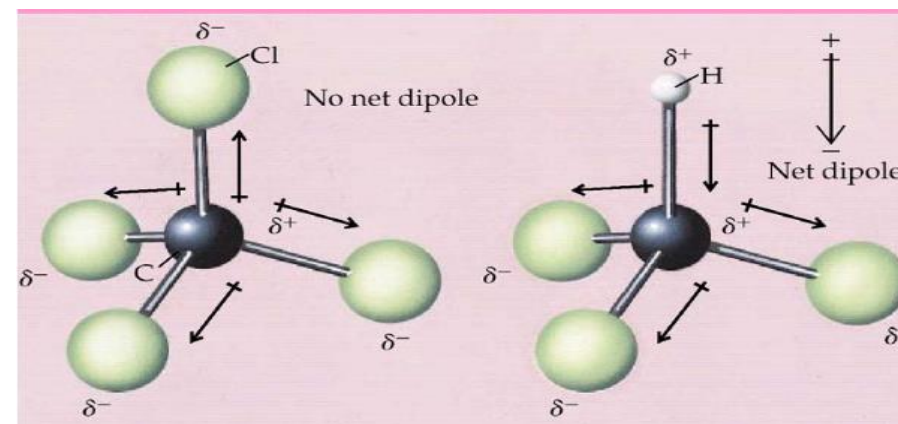
- ❖ Use electronegativity values to predict bond dipoles.
- ❖ Use the VSEPR method to predict the molecular shape.
- ❖ From the molecular shape, determine whether bond dipoles cancel to give a nonpolar molecule, or combine to produce a **resultant dipole moment for the molecule!!**

**Note:** Lone-pair electrons can also make a contribution to dipole moments.

# Polar or non-polar?

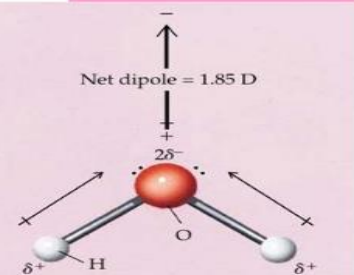


Which molecule is polar? You must look at the geometry to decide.



$\text{NH}_3$

$\text{H}_2\text{O}$



**CONCEPT CHECK!****True or false:**

A molecule that has polar bonds will always be polar.

- If true, explain why.
- If false, provide a counter-example

**True or false:**

Lone pairs make a molecule polar.

- If true, explain why.
- If false, provide a counter-example.