ملخص الوحدة الثالثة The Covalent Bond الرابطة التساهمية منهج انسباير

The covalent bond

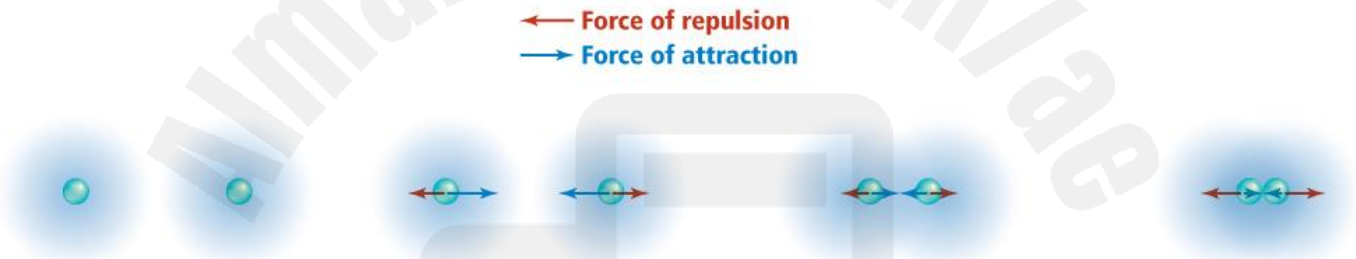
Section 1 :

Why do atoms bond?

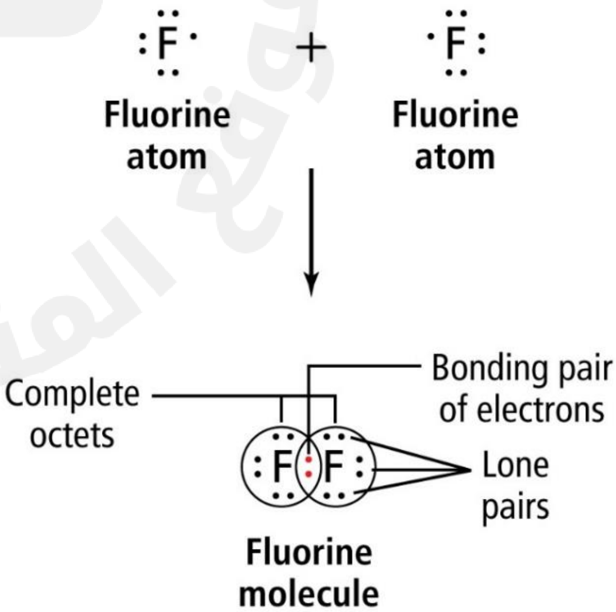
• Atoms gain stability when they share electrons and form covalent bonds.
• Lower energy states make an atom more stable.
• Gaining or losing electrons makes atoms more stable by forming ions with noble-gas electron configurations.
• Sharing valence electrons with other atoms also results in noble-gas electron configurations.

• Atoms in non-ionic compounds share electrons.
• The chemical bond that results from sharing electrons is a **covalent bond**.
• A **molecule** is formed when two or more atoms bond.

• Diatomic molecules (H_2, F_2 for example) exist because two-atom molecules are more stable than single atoms.



The most stable arrangement of atoms exists at the point of maximum net attraction, where the atoms bond covalently and form a molecule.



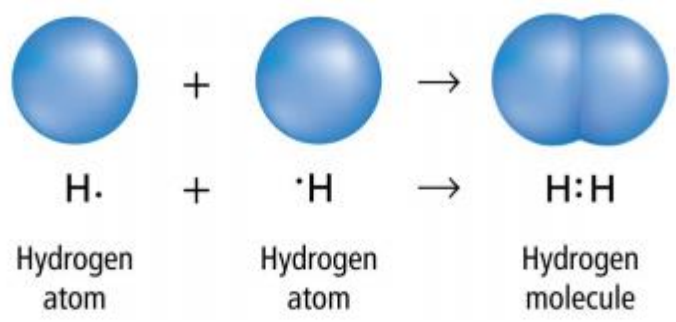
Covalent bonds are different from ionic bonds because:

A. atoms in a covalent bond lose electrons to another atom
B. atoms in a covalent bond do not have noble-gas electron configurations
C. atoms in a covalent bond share electrons with another atom
D. atoms in covalent bonds gain electrons from another atom

Single Covalent Bonds

• When only one pair of electrons is shared, the result is a **single** covalent bond.

• The figure shows two hydrogen atoms forming a hydrogen molecule with a single covalent bond, resulting in an electron configuration like helium **(lower energy).**



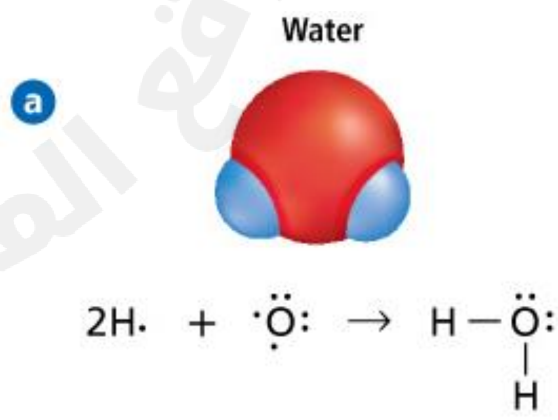
(**H** can form **only** single covalent bond)

• In a **Lewis structure** dots or a line are used to symbolize a single covalent bond.

• The halogens—the group 17 elements—have 7 valence electrons and form single covalent bonds with atoms of other non-metals.

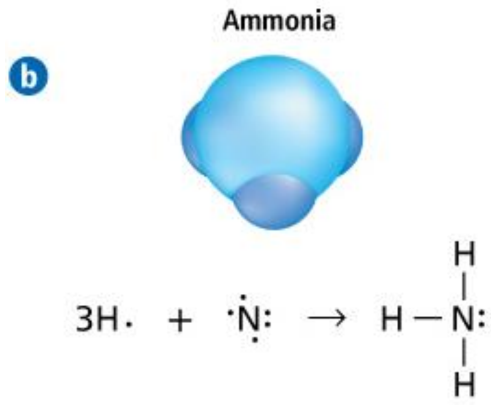
• Atoms in group 16 can share two electrons and form two covalent bonds.

• Water is formed from one oxygen with two hydrogen atoms covalently bonded to it.



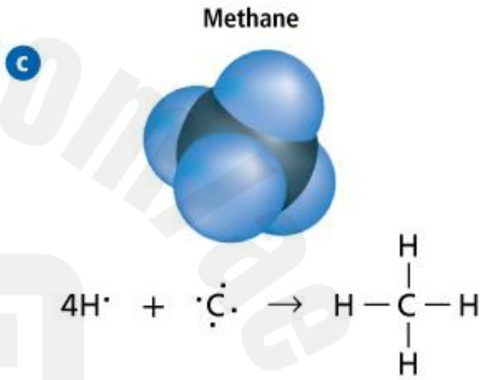
Two Single Covalent Bonds

- Atoms in group 15 form three single covalent bonds, such as in ammonia.



Three Single Covalent Bonds

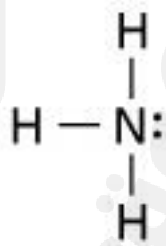
- Atoms of group 14 elements form four single covalent bonds, such as in methane.



Four Single Covalent Bonds

Application:

1. Where are bonding pairs and lone pairs?



2. Draw the Lewis structure for each molecule.

HF, PH_3, H_2S, HCl, CCl_4, SiH_4, NF_3

• **Sigma bonds (Ơ)** are single covalent bonds.

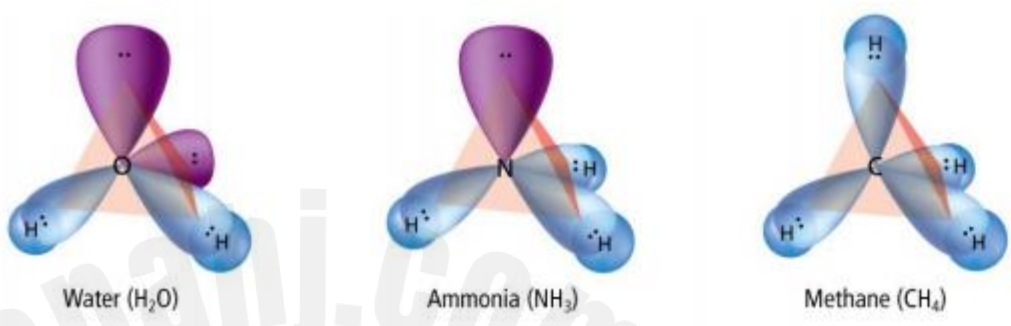
• Sigma bonds occur when the pair of shared electrons is in an area centered between the two atoms.

In the following cases:

s orbital overlaps with another s orbital

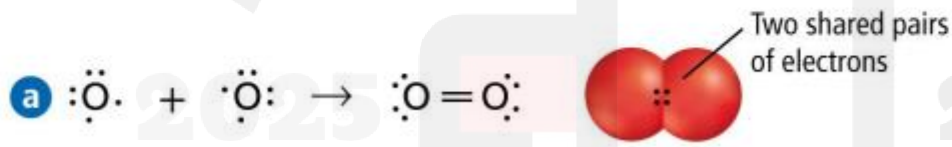
s orbital overlaps with another p orbital

two p orbitals overlap end-to-end.

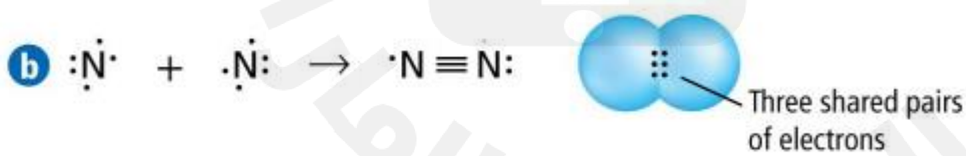


Multiple Covalent Bonds

• Double bonds form when two pairs of electrons are shared between two atoms.

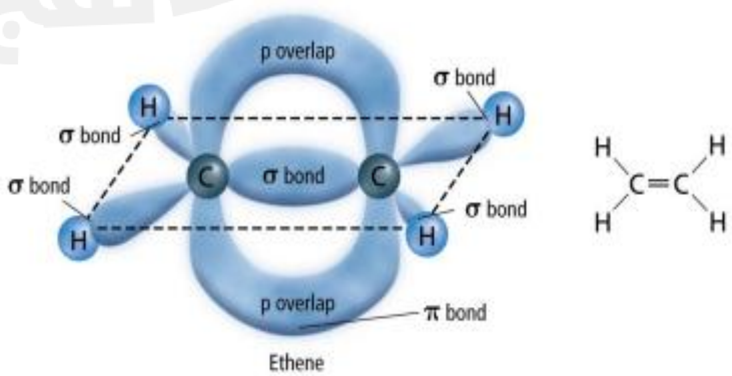


• Triple bonds form when three pairs of electrons are shared between two atoms.



• A multiple covalent bond consists of one sigma bond and at least one pi bond.

• The **pi bond (π)** is formed when parallel orbitals overlap and share electrons.



What does a triple bond consist of?

A. three sigma bonds

B. three pi bonds

C. two sigma bonds and one pi bond

D. two pi bonds and one sigma bond

The Strength of Covalent Bonds

•The strength depends on the distance between the two nuclei, or **bond length**.

• As length increases, strength decreases.

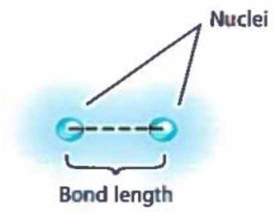


Table 8.1 Covalent Bond Type and Bond Length

Molecule	Bond Type	Bond Length
F_2	single covalent	1.43×10^{-10} m
O_2	double covalent	1.21×10^{-10} m
N_2	triple covalent	1.10×10^{-10} m

• The amount of energy required to break a bond is called **the bond dissociation energy.**

• The shorter the bond length, the greater the energy required to break it.

Table 8.2 Bond-Dissociation Energy

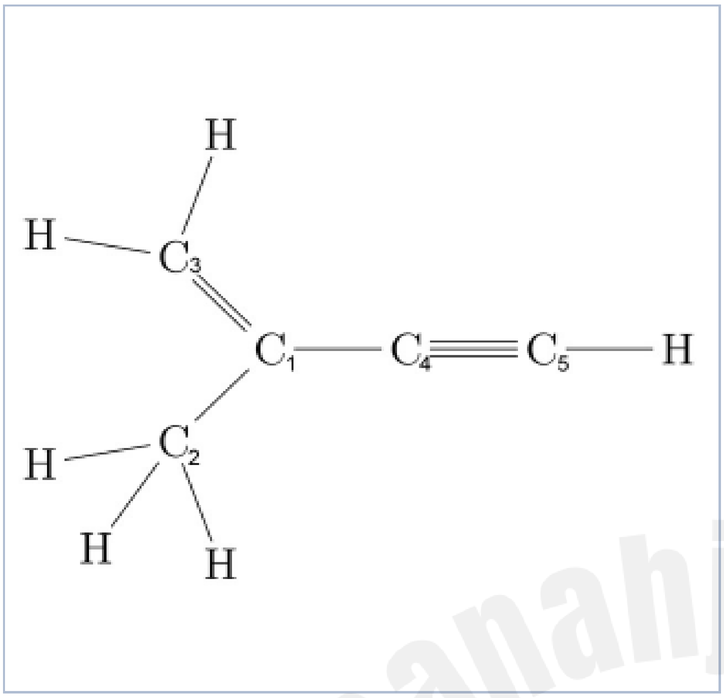
Molecule	Bond-Dissociation Energy
F_2	159 kJ/mol
O_2	498 kJ/mol
N_2	945 kJ/mol

• An **endothermic reaction** is one where a greater amount of energy is required to break a bond in reactants than is released when the new bonds form in the products.

• An **exothermic reaction** is one where more energy is released than is required to break the bonds in the initial reactants.

Determine all sigma and pi bonds on the previous figure.

Activity: How many bonds does each carbon atom form in the molecule?



Carbon Atoms	Single Bond	Double Bond	Triple Bond
C1			
C2			
C3			
C4			
C5			

The strongest bond is:

A. Between C4 and C5
B. Between C1 and C2
C. Between C1 and C4
D. Between C2 and H

Section 2 Naming Molecules

Table 8.3	Prefixes in Covalent Compounds		
Number of Atoms	Prefix	Number of Atoms	Prefix
1	mono-	6	hexa-
2	di-	7	hepta-
3	tri-	8	octa-
4	tetra-	9	nona-
5	penta-	10	deca-

Naming binary molecular compound

[#of atoms] Same name / [#of atoms] root + id name

Example:

N_2O P_2O

CO NO_2

CO_2 SO_2

NF_3 CCl_4

What is the name of: O_2F_2?

a. dioxide difluoride
b. dioxide difluoride
c. dioxygen difluoring
d. dioxygen difluride

What is the name of: BrCl?

a. monbromide monchloride
b. bromine monchloride
c. bromide monchloride

Writing formula

Example:

Dinitrogen monoxide

Write the formula of the following compounds:

dihydrogen oxide chlorine trifluoride diphosphorus trioxide

dinitrogen trioxide nitrogen monoxide

What is the correct formula of tetraphosphorus decaoxide?

P_2O_{10} P_4O_{10} $P_{10}O_5$

Naming acids

Acid: is a compound H/ anion

Binary acid (gases form acids when dissolve in water)

H X

Hydro-X-ic acid

Example:

HCl ……………………………………………

HBr……………………………………………

HF ……………………………………………

HI……………………………………………

Formula of Ion	Name of Ion
Nonmetals	
S^{2-}	Sulfide
F^-	Fluoride
Cl^-	Chloride
Br^-	Bromide
I^-	Iodide

Challenge:

H_2S ……………………………………………

HCN ……………………………………………

Oxyacid:

H X O

X-ate X-ic acid

X-ite X-ous acid

Example:

$HClO_3$ ....

$HClO_2$ ....

Polyatomic ions

Ion	Name	Ion	Name	Ion	Name
NO_3^-	nitr***ate***	PO_4^{3-}	phosph***ate***	$C_2H_3O_2^-$	acet***ate***
NO_2^-	nitr***ite***	HPO_4^{2-}	Hydrogen phosph***ate***	AsO_4^{3-}	arsen***ate***
SO_4^{2-}	sulf***ate***	$H_2PO_4^-$	dihydrogen phosph***ate***	MnO_4^-	permangan***ate***
SO_3^{2-}	sulf***ite***	CO_3^{2-}	cabon***ate***		
$S_2O_3^{2-}$	thiosulf***ate***	HCO_3^-	hydrogen carbon***ate***	OH^-	hydroxide
		CrO_4^{2-}	chrom***ate***	CN^-	cyanide
		$Cr_2O_7^{2-}$	dichrom***ate***	NH_4^+	ammonium

Ion	Name	Ion	Name	Ion	Name
ClO_4^-	***per***chlor***ate***			IO_4^-	***per***iod***ate***
ClO_3^-	chlor***ate***	BrO_3^-	brom***ate***	IO_3^-	Iod***ate***
ClO_2^-	chlor***ite***				
ClO^-	***hypo***chlor***ite***				

Name the following acids:

HNO_3 .... HNO_2 ....

H_2CO_3 .... $HBrO_3$ ....

H_2SO_4 .... H_2SO_3 ....

H_3PO_4 .... H_3PO_3 ....

$HClO_4$ .... $HClO$ ....

Match: Column A **Column B**

1. H_2CO_3 **a.** hydrobromic acid

2. HNO_2 **b.** nitrous acid

3. HNO_3 **c.** nitric acid

4. HBr **d.** carbonic acid

5. $HBrO_3$ **e.** bromic acid

• **Many compounds** were discovered and given common names long before the present naming system was developed (ie. water, ammonia, hydrazine)

Table 8.5	Formulas and Names of Some Covalent Compounds	Interactive Table Explore naming covalent compounds glencoe.com.
Formula	Common Name	Molecular Compound Name
H_2O	water	dihydrogen monoxide
NH_3	ammonia	nitrogen trihydride
N_2H_4	hydrazine	dinitrogen tetrahydride
HCl	muriatic acid	hydrochloric acid
$C_9H_8O_4$	aspirin	2-(acetyloxy)benzoic acid

• The name of a molecular compound reveals its composition and is important in communicating the nature of the compound.

Extra: Name the following acids.

HI ……………………………… $HClO_3$ ………………………………

$HClO_2$ ……………………………… H_2SO_4………………………………

H_2S ………………………………

Give the formula for each compound.

dihydrogen monoxide

chorine trifluoride

diphosphorus trioxide

disulfur decafluoride

dinitrogen trioxide

nitrogen monoxide

hydrochloric acid

chloric acid

sulfuric acid

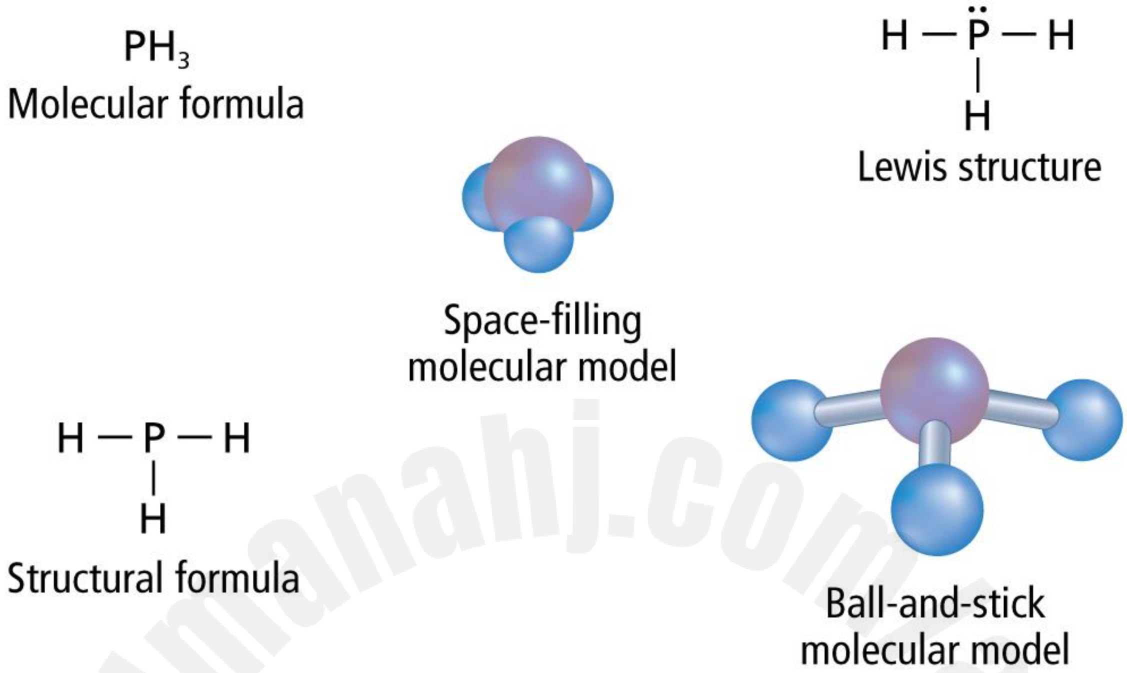
sulfurous acid

carbonic acid

periodic acid

Section 3: Molecular Structures

A **structural formula** uses letter symbols and bonds to show relative positions of
atoms. (the most useful molecular model)



Drawing Lewis Structures

1. – Predict the location of certain atoms. (the central atom is usually the one closer to left side in periodic table). Other atoms are terminal (end) atoms. **H is always a terminal atom**.
2. – Determine the number of electrons available for bonding. (total Valence electrons)
3. – Determine the number of bonding pairs. (total Valence electrons/2)
4. – Place the bonding pairs. (single bond between central atom and each terminal)
5. – Determine the number of bonding pairs remaining. (total pairs – used pairs)

   place lone pairs around each terminal atom to satisfy the octet rule.

6. – Any remaining pairs will be assigned to the central atom.

Determine whether the central atom satisfies the octet rule. If not, convert one or two lone pairs on the terminal atom into a double or triple bond between the terminal atom and the central atom.

(C, N, O, S often form double or triple bonds).

Note: Atoms within a polyatomic ion are covalently bonded.

Exercises:

Draw the Lewis structure for **(PH_3)**

Draw the Lewis structure for ammonia (NH_3)

Draw the Lewis structure for **(NF_3)**

Draw the Lewis structure for carbon dioxide **(CO_2)**

Draw the Lewis structure for carbon disulfide (CS_2)

Draw the Lewis structure for ethylene (C_2H_2)

Lewis structure for polyatomic ions:

Draw the Lewis structure for phosphate **(PO_4^{3-})**

Draw the Lewis structure for **(NH_4^+)**

Draw the Lewis structure for Chlorate (ClO_4^-)

Resonance Structures

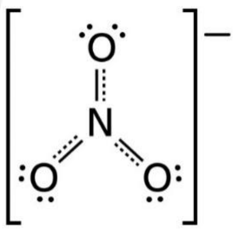
• **Resonance** is a condition that occurs when more than one valid Lewis structure can be written for a molecule or ion.

When a molecule or polyatomic ion has both a double bond and a single bond.

Resonance structures differs only in the position of the electron pairs (lone and bonding pairs), never the atom positions.

• This figure shows **three correct ways** to draw the structure for $(NO_3)^-$.

• Two or more correct Lewis structures that represent a single ion or molecule are resonance structures.
• The molecule behaves as though it has only one structure.
• The bond lengths are identical to each other and intermediate between single and double covalent bonds.



The actual length is an average of the bonds in the resonance structures.

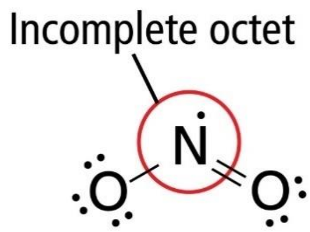
Examples of resonance:
Draw the Lewis resonance structure for the following molecules.

O_3, NO_3^-, **NO_2^-,** SO_3^{2-}, CO_3^{2-}, **SO_2**

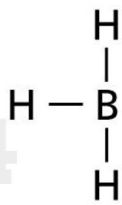
Exceptions to the Octet Rule

• Some molecules do not obey the octet rule.

1. **Odd number of valence electrons.**
   • A small group of molecules might have an odd number of valence electrons.
   • NO_2 has five valence electrons from nitrogen and 12 from oxygen and cannot form an exact number of electron pairs.
   Examples: ClO_2, **NO**



2. **Suboctete covalent bonds:**
   A few compounds form stable configurations with **less than 8** electrons around the atom. <u>(central atom is from group 2 or 13)</u>
   Examples: **<u>BH_3</u>**

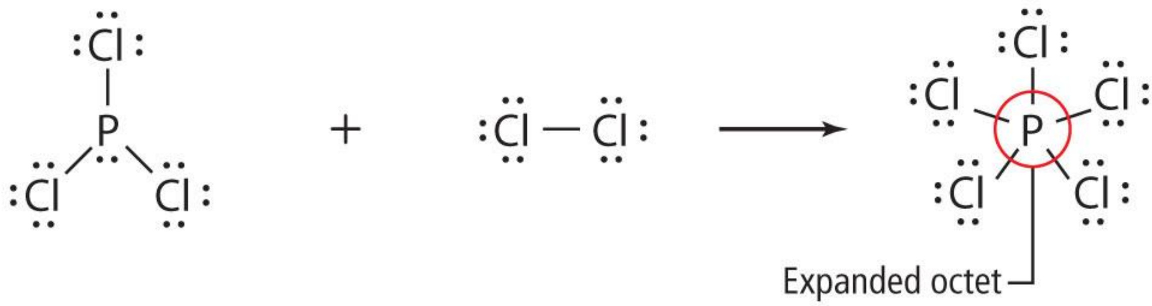


A **coordinate covalent bond** forms when one atom donates both of the electrons to be shared with an atom or ion that needs two electrons.

The boron atom has no electrons to share, whereas the nitrogen atom has two electrons to share.

The nitrogen atom shares both electrons to form the coordinate covalent bond.

3. **Expanded octet:**
   - A third group of compounds has central atoms with more than eight valence electrons, called an expanded octet.
   - Elements in period 3 or higher have a d-orbital and can form more than four covalent bonds.



Examples:

Draw the Lewis structure for the following molecules.

XeF_4

ClF_3

PCl_5

SF_6

Draw the Lewis structure for the following molecules.

SiF_4, CN^-, HCO_3^-, AsF_6^-

Draw the Lewis resonance structure for dinitrogen monoxide (N_2O).

Section 4 Molecular Shapes

VSEPR Model

• The shape of a molecule determines many of its physical and chemical properties.

• Molecular geometry (shape) can be determined with the Valence Shell Electron Pair Repulsion model, or **VSEPR model** which minimizes the repulsion of shared and unshared atoms around the central atom.

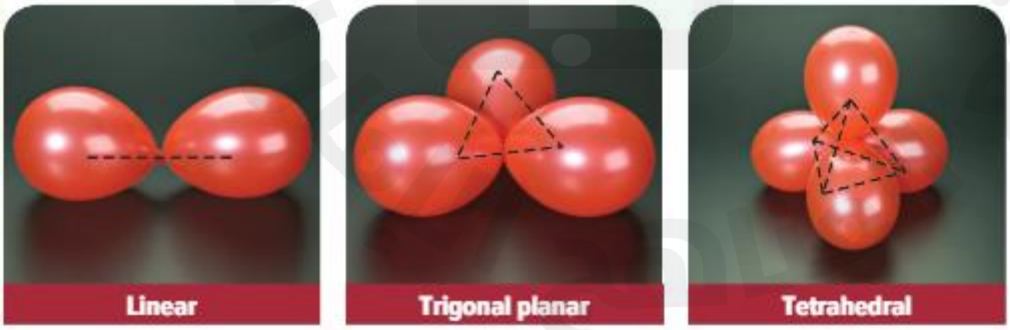
<u>Bond angle:</u>

• Electron pairs repel each other and cause molecules to be in fixed positions relative to each other.

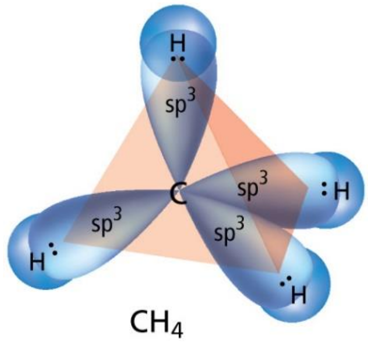
• Unshared electron pairs also determine the shape of a molecule.

• Electron pairs are located in a molecule as far apart as they can be.

•Unshared pairs of electrons occupy a larger orbital than shared pairs

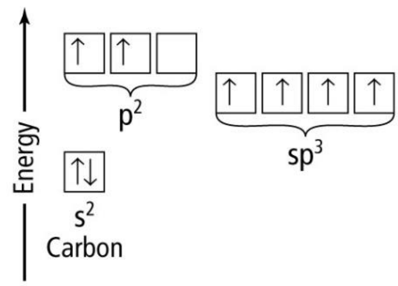


Hybridization is a process in which atomic orbitals mix and form new, identical hybrid orbitals.



• Carbon often undergoes hybridization, which forms an sp^3 orbital formed from one s orbital and three p orbitals.

Carbon has four sp^3 orbitals



• Lone pairs also occupy hybrid orbitals.

• Single, double, and triple bonds occupy only **one** hybrid orbital (CO_2 with two double bonds forms an sp hybrid orbital).

$AlCl_3$

has three pairs of electrons:

VSPER predicts a **trigonal planar** molecular shape

<u>(**Al**)</u> has **three** identical **sp2** hybrid orbitals.

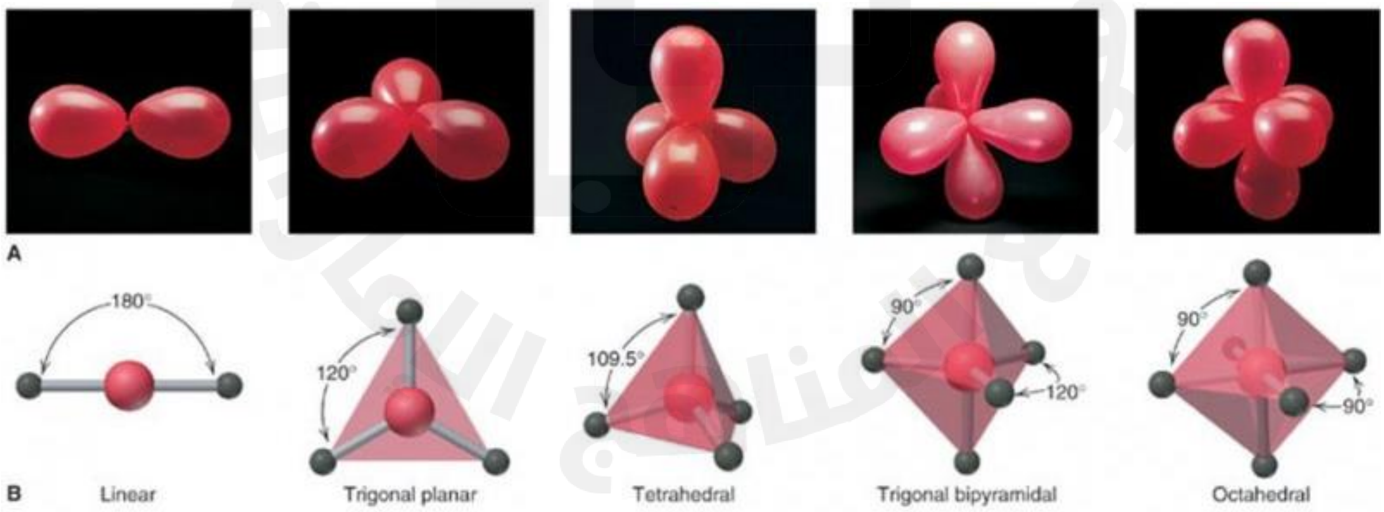
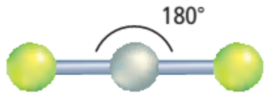
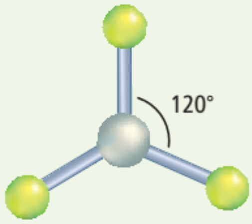
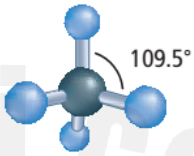
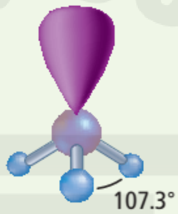
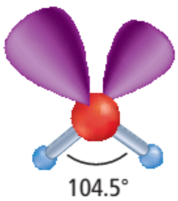
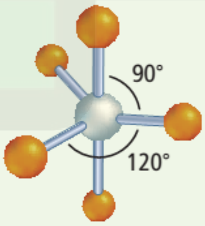
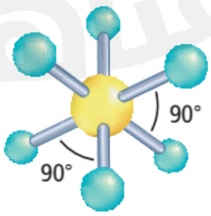


Table Molecular Shapes

Molecule	Total Pairs	Shared Pairs	Lone Pairs	Hybrid Orbitals	Molecular Shape*
$BeCl_2$	2	2	0	sp	 Linear
$AlCl_3$	3	3	0	sp^2	 Trigonal planar
CH_4	4	4	0	sp^3	 Tetrahedral
PH_3	4	3	1	sp^3	 Trigonal pyramidal
H_2O	4	2	2	sp^3	 Bent
$NbBr_5$	5	5	0	sp^3d	 Trigonal bipyramidal
SF_6	6	6	0	sp^3d^2	 Octahedral

*Balls represent atoms, sticks represent bonds, and lobes represent lone pairs of electrons.

The $BeCl_2$ molecule contains only two pairs of electrons shared with the central Be atom. These bonding electrons have the maximum separation, a bond angle of 180°, and the molecular shape is linear.

The three bonding electron pairs in $AlCl_3$ have maximum separation in a trigonal planar shape with 120° bond angles.

When the central atom in a molecule has four pairs of bonding electrons, as CH_4 does, the shape is tetrahedral. The bond angles are 109.5°.

PH_3 has three single covalent bonds and one lone pair. The lone pair takes up a greater amount of space than the shared pairs. There is stronger repulsion between the lone pair and the bonding pairs than between two bonding pairs. The resulting geometry is trigonal pyramidal, with 107.3° bond angles.

Water has two covalent bonds and two lone pairs. Repulsion between the lone pairs causes the angle to be 104.5°, less than both tetrahedral and trigonal pyramidal. As a result, water molecules have a bent shape.

The $NbBr_5$ molecule has five pairs of bonding electrons. The trigonal bipyramidal shape minimizes the repulsion of these shared electron pairs.

As with $NbBr_5$, SF_6 has no unshared electron pairs on the central atom. However, six shared pairs arranged about the central atom result in an octahedral shape.

Application:

Determine the molecular shape, bond angle, and hybrid orbitals for (PH_3)

The molecular shape is trigonal pyramidal with a predicted 107° bond angle and sp^3 hybrid orbitals.

Determine the Lewis structure, molecular shape, bond angle, and hybrid orbitals for each molecule.

BF_3 , OCl_2 , BeF_2 , CF_4 , $NH4^+$,

CS_2 , CH_2O , H_2Se , CCl_2F_2 , NCl_3.

Section 5 Electronegativity and Polarity

Electron Affinity, Electronegativity, and Bond Character

• **Electron affinity** measures the tendency of an atom to accept an electron.

Electronegativity Values for Selected Elements

Metal / Metalloid / Nonmetal

1 H 2.20																
3 Li 0.98	4 Be 1.57											5 B 2.04	6 C 2.55	7 N 3.04	8 O 3.44	9 F 3.98
11 Na 0.93	12 Mg 1.31											13 Al 1.61	14 Si 1.90	15 P 2.19	16 S 2.58	17 Cl 3.16
19 K 0.82	20 Ca 1.00	21 Sc 1.36	22 Ti 1.54	23 V 1.63	24 Cr 1.66	25 Mn 1.55	26 Fe 1.83	27 Co 1.88	28 Ni 1.91	29 Cu 1.90	30 Zn 1.65	31 Ga 1.81	32 Ge 2.01	33 As 2.18	34 Se 2.55	35 Br 2.96
37 Rb 0.82	38 Sr 0.95	39 Y 1.22	40 Zr 1.33	41 Nb 1.6	42 Mo 2.16	43 Tc 2.10	44 Ru 2.2	45 Rh 2.28	46 Pd 2.20	47 Ag 1.93	48 Cd 1.69	49 In 1.78	50 Sn 1.96	51 Sb 2.05	52 Te 2.1	53 I 2.66
55 Cs 0.79	56 Ba 0.89	57 La 1.10	72 Hf 1.3	73 Ta 1.5	74 W 1.7	75 Re 1.9	76 Os 2.2	77 Ir 2.2	78 Pt 2.2	79 Au 2.4	80 Hg 1.9	81 Tl 1.8	82 Pb 1.8	83 Bi 1.9	84 Po 2.0	85 At 2.2
87 Fr 0.7	88 Ra 0.9	89 Ac 1.1														

• Noble gases are not listed because they generally do not form compounds.

Electron affinity increases in period left to right but decreases in group up to down.

Electronegativity: the relative ability of an atom to attract electrons in a chemical bond

This table lists the character and type of chemical bond that forms with differences in electronegativity.

Table 8.7 EN Difference and Bond Character

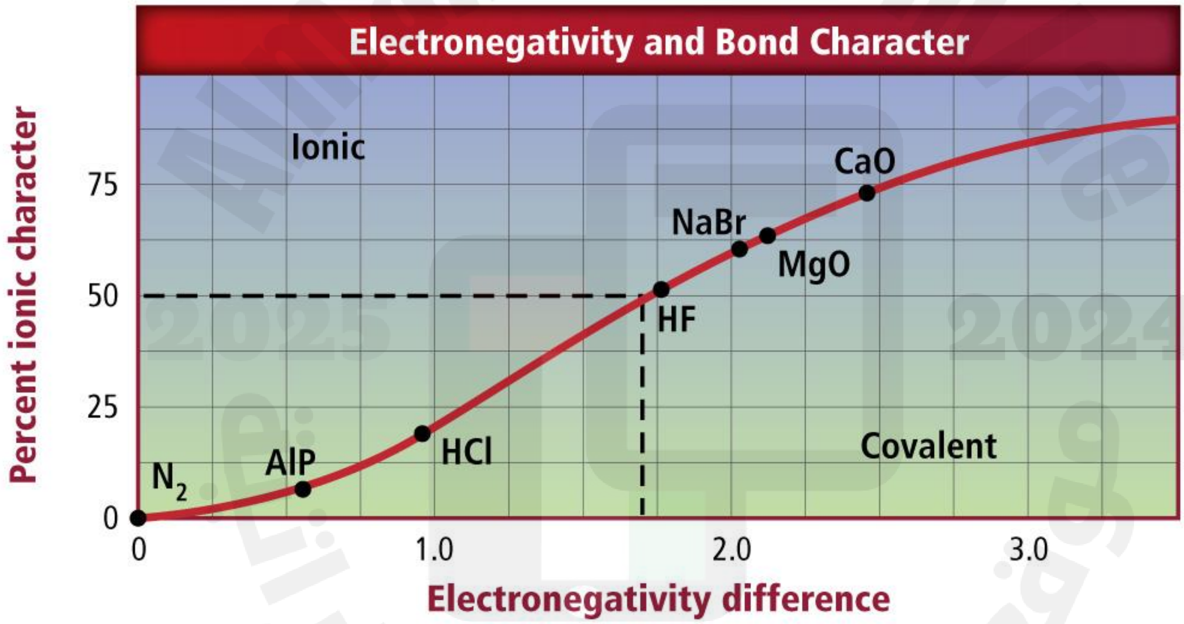
Electronegativity Difference	Bond Character
> 1.7	mostly ionic
0.4 – 1.7	polar covalent
< 0.4	mostly covalent
0	nonpolar covalent

• Unequal sharing of electrons results in a **polar covalent bond**.

• Bonding is often not clearly ionic or covalent.

F (fluorine) has the **highest** electronegativity and **Fr (francium)** the **least** electronegativity

This graph summarizes the range of chemical bonds between two atoms.



What percent ionic character is a bond between two atoms that have an electronegativity difference of 2.00? ……………………………………………………..

Where would LiBr be plotted on the graph? ……………………………………………..

……

What is the percent ionic character of a pure covalent bond? **……………………**

Determine the percent ionic character of calcium oxide. ………………………….

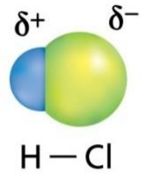
……

……

Polar Covalent Bonds: form when atoms pull on electrons in a molecule <u>unequally</u>.

- Electrons spend more time around one atom than another resulting in partial charges at the ends of the bond called a dipole.

Electronegativity Cl = 3.16
Electronegativity H = 2.20
Difference = 0.96



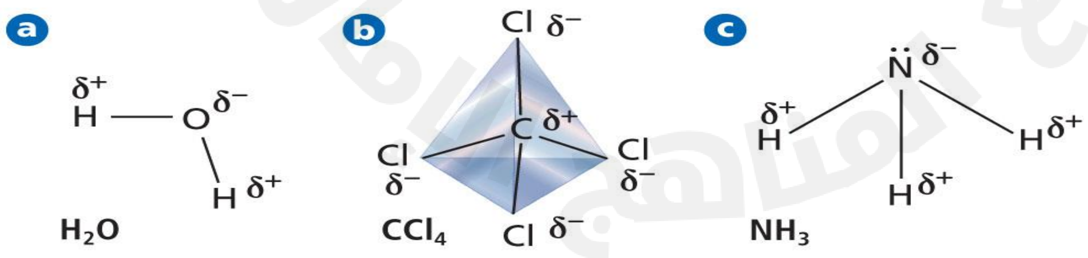
* **Delta (δ)** is used to represent a partial charge. In a polar covalent bond, δ^- represents a partial negative charge and δ^+ represents a partial positive charge. And can be added to a molecular model to indicate the polarity of the covalent bond. The resulting polar bond often is referred to as a dipole (two poles)

Molecular polarity: Covalently bonded molecules are either polar or non-polar.

- Non-polar molecules are not attracted by an electric field.
- Polar molecules align with an electric field.

Compare water (H_2O) and CCl_4. H — O 1.24, C — Cl 0.61

- Both bonds are polar, but only water is a polar molecule because of the shape of the molecule so the charge distribution in unequal.
- The electric charge on a CCl_4 molecule measured at any distance from the center of the molecule is identical to the charge measured at the same distance on the opposite side.



The bent shape of a water molecule makes it polar.

The symmetry of a CCl_4 molecule results in an equal distribution of charge, and the molecule is nonpolar.

The asymmetric shape of an ammonia molecule results in an unequal charge distribution and the molecule is polar.

Is NH_3 polar? Explain. ……………………………………………………………………………….

• Solubility is the property of a substance's ability to dissolve in another substance.

• Polar molecules and ionic substances are usually soluble in polar substances.

• Non-polar molecules dissolve only in nonpolar substances.

I.e. Water and oil cannot be mixed. Water alone will not clean oil from a fabric.

Properties of Covalent Compounds

• Covalent bonds between atoms are strong, but attraction forces between molecules are weak.

• The weak attraction forces are known as van der Waals forces.

• The forces vary in strength but are weaker than the bonds in a molecule or ions in an ionic compound.

1) Non-polar molecules exhibit a weak **dispersion force**, or induced dipole.
2) **Dipole-dipole force** is the force between two oppositely charged ends of two polar molecules.
3) **A hydrogen bond** is an especially strong dipole-dipole force between a hydrogen end of one dipole and a fluorine, oxygen, or nitrogen atom on another dipole.

• Many physical properties are due to intermolecular forces.

• Weak forces result in the relatively low melting and boiling points of molecular substances.

Salt does not melt but sugar melts at a relatively low temperature.

Example: O_2, CO_2 , H_2S are gases at room temperature?

• Many covalent molecules are relatively soft solids.

Example: Paraffin (found in candles) is solid at room temperature.

• Molecules can align in a crystal lattice, similar to ionic solids but with less attraction between particles.

• Solids composed of only atoms interconnected by a network of covalent bonds are called covalent network solids.

• Quartz and diamonds are two common examples of **network solids.**