Chemical Bonding

Chemical Bonds

1. Ionic bonds
2. Covalent bonds
3. Metallic bonds
4. Hydrogen bonds
5. Van der Waals forces

Chemical Bonds

1. Ionic bonds

2. Covalent bonds

Review

main-group elements

1A – 8A

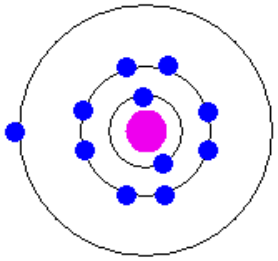
Level 1 → Maximum 2 electrons in valence level

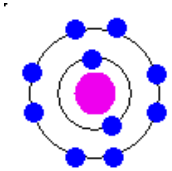
Hydrogen and Helium

Other levels → Maximum 8 electrons in valence level

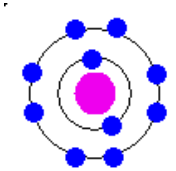
Octet & Duet rules

Goal of atoms ⟶ Filled valence level ⟶ Noble gases (Stable)

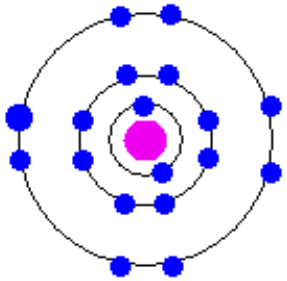




+ e^-



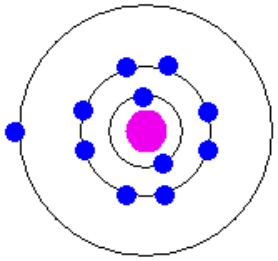
Na: $1s^2\ 2s^2\ 2p^6\ 3s^1$ Na^+: $1s^2\ 2s^2\ 2p^6$ Ne: $1s^2\ 2s^2\ 2p^6$



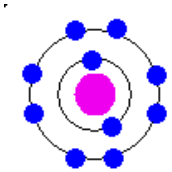
Ar: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$

Octet & Duet rules

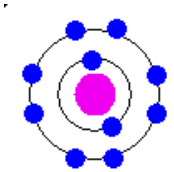
Goal of atoms ⟶ Filled valence level ⟶ Noble gases (Stable)



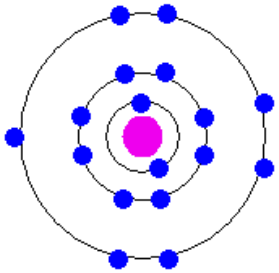
⟶



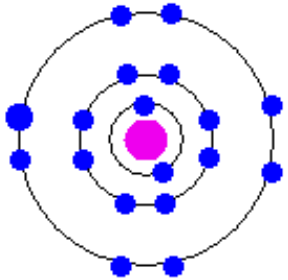
+ e^-

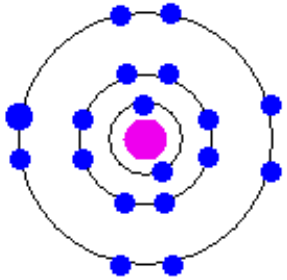


Na: $1s^2\ 2s^2\ 2p^6\ 3s^1$ Na^+: $1s^2\ 2s^2\ 2p^6$ Ne: $1s^2\ 2s^2\ 2p^6$



+ e^- ⟶

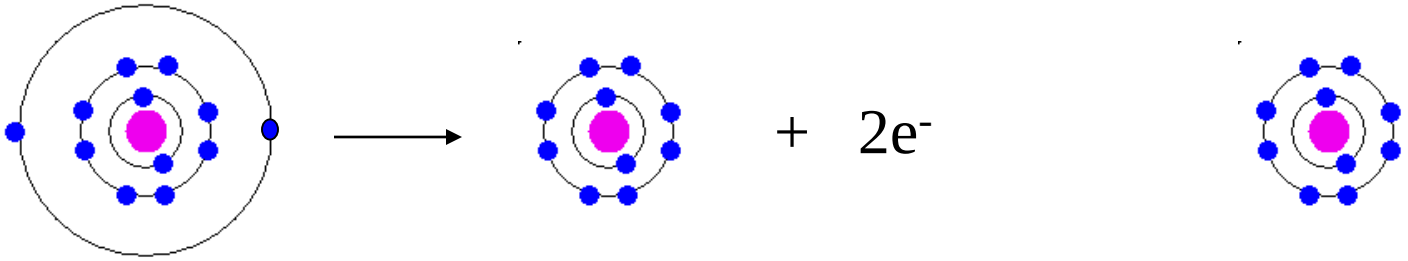




Cl: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$ Cl^-: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$ Ar: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$

Octet & Duet rules

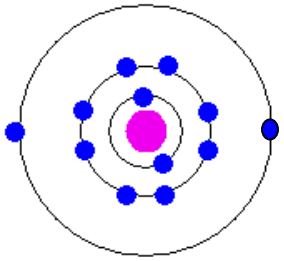
Goal of atoms ⟶ Filled valence level ⟶ Noble gases (Stable)

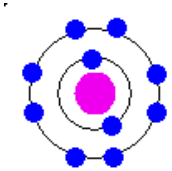


Mg: $1s^2\ 2s^2\ 2p^6\ 3s^2$ Mg^{2+}: $1s^2\ 2s^2\ 2p^6$ Ne: $1s^2\ 2s^2\ 2p^6$

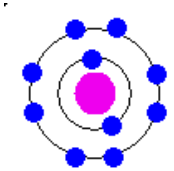
Octet & Duet rules

Goal of atoms ⟶ Filled valence level ⟶ Noble gases (Stable)

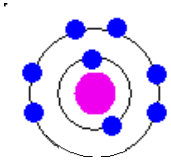




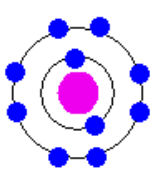
+ $2e^-$

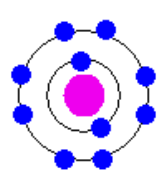


Mg: $1s^2\ 2s^2\ 2p^6\ 3s^2$ Mg^{2+}: $1s^2\ 2s^2\ 2p^6$ Ne: $1s^2\ 2s^2\ 2p^6$



+ $2e^-$ ⟶





O: $1s^2\ 2s^2\ 2p^4$ O^{2-}: $1s^2\ 2s^2\ 2p^6$ Ne: $1s^2\ 2s^2\ 2p^6$

Metals: lose 1, 2 or 3 e^- → Cation (Y^+)

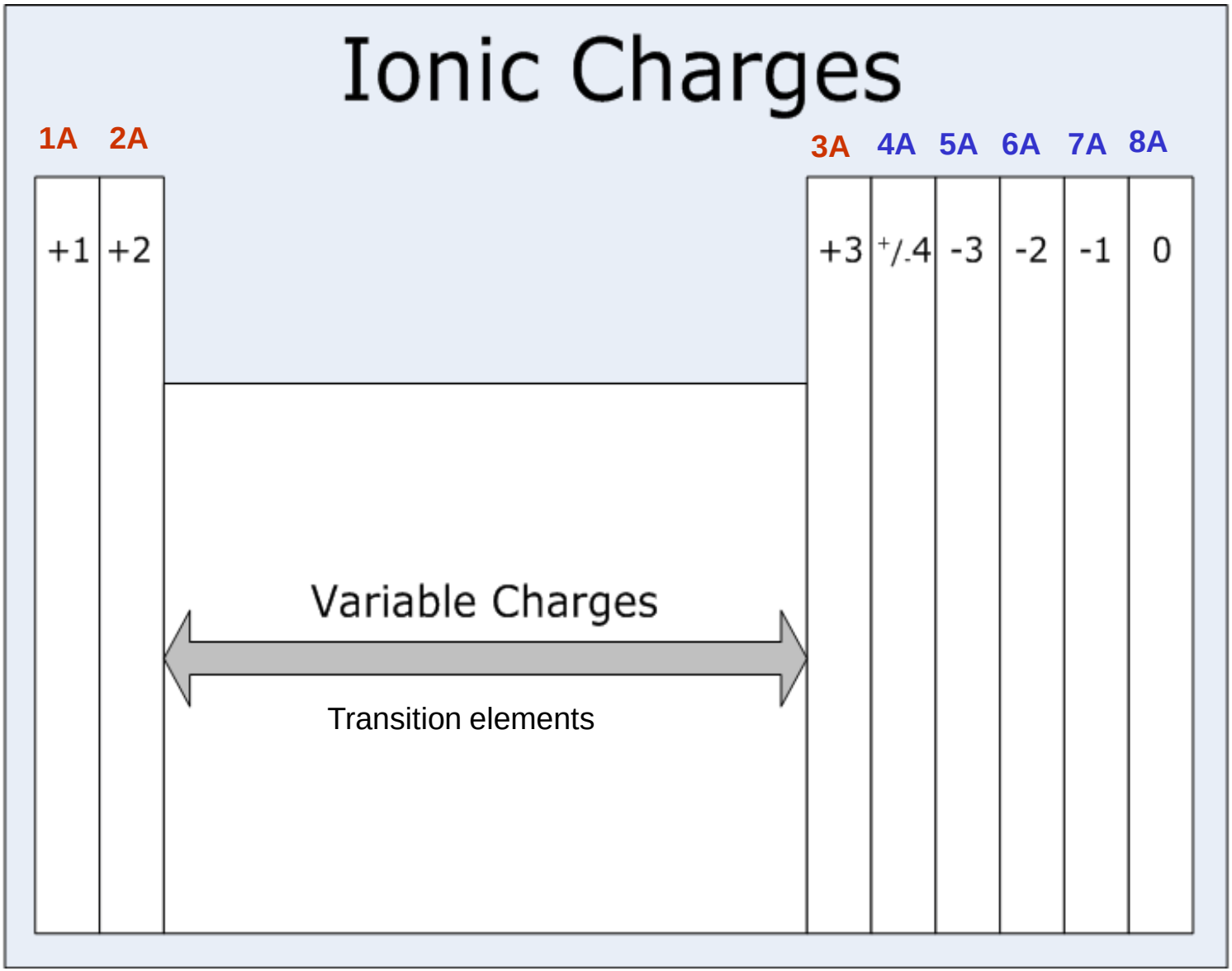
Nonmetals: gain 1, 2 or 3 e^- → Anion (X^-)

} Ions

Number of protons and neutrons in the nucleus remains unchanged.

Cation (Y^+): Na^+ Li^+ Ca^{2+} Al^{3+}

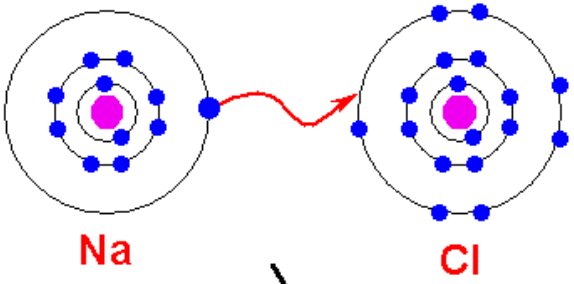
Anion (X^-): Cl^- F^- O^{-2}


Ionic Charges
1A
2A
3A
4A
5A
6A
7A
8A
+1
+2
+3
+/-4
-3
-2
-1
0
Variable Charges
Transition elements

Ionic bonds

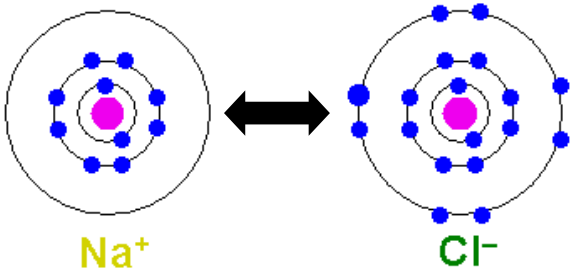
Metal-Nonmetal

Na: $1s^2\ 2s^2\ 2p^6\ 3s^1$



Cl: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$

Na^+: $1s^2\ 2s^2\ 2p^6$

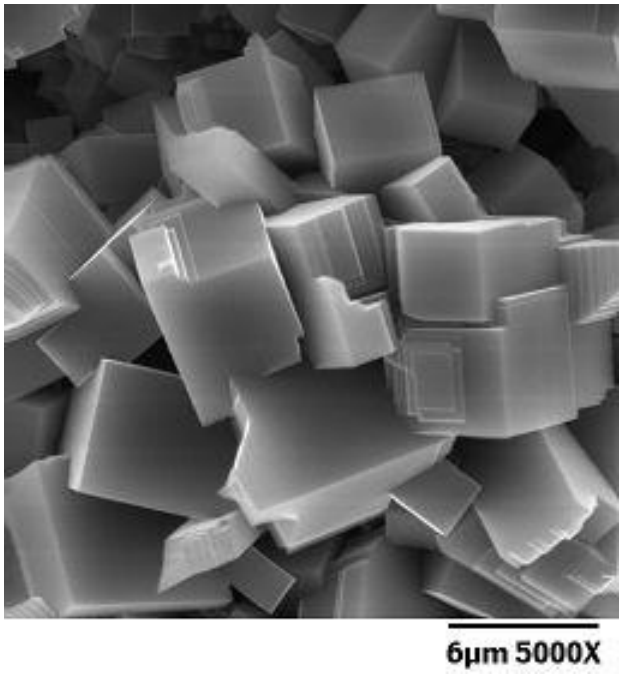


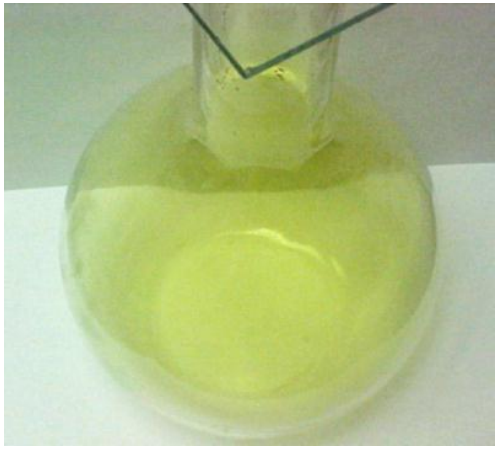
Cl^-: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$

Cation

Anion

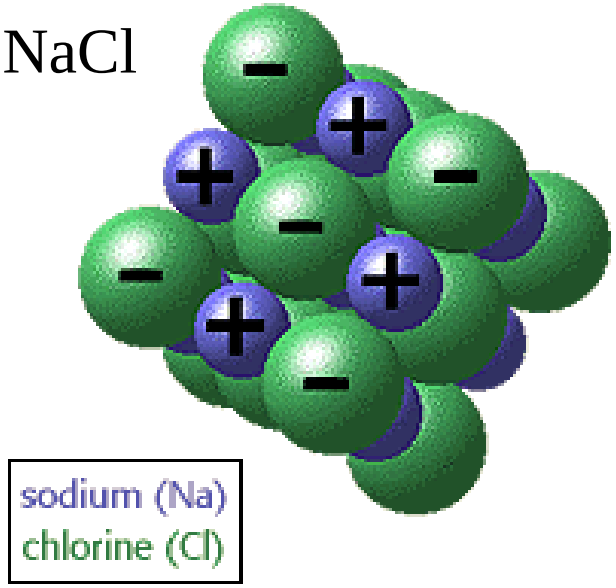
Sodium (Na)





Chlorine (Cl)

NaCl



matter are neutral (uncharged):

total number of positive charges = total number of negative charges

Na^{+} Cl^{-} ⟶ $NaCl$

Ca^{2+} Cl^{-} ⟶ $CaCl_2$

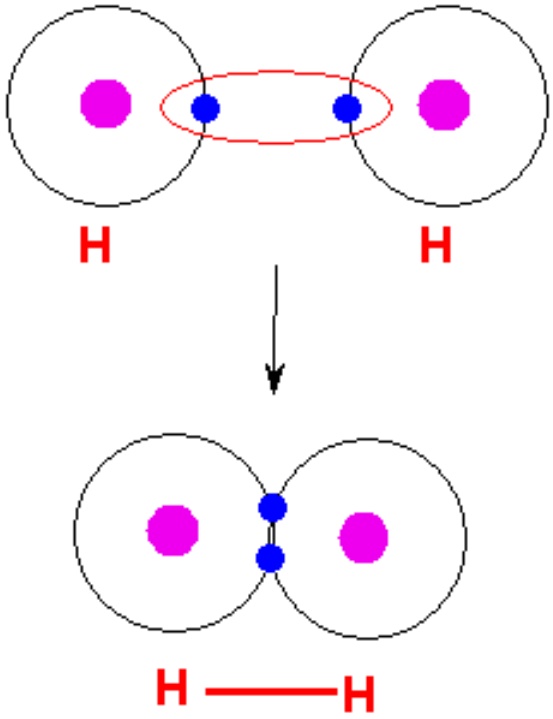
Al^{3+} S^{2-} ⟶ Al_2S_3

Ba^{2+} O^{2-} ⟶ ~~Ba_2O_2~~ BaO

Covalent bonds

Nonmetal-Nonmetal
Metalloid-Nonmetal

Sharing of
valence electrons

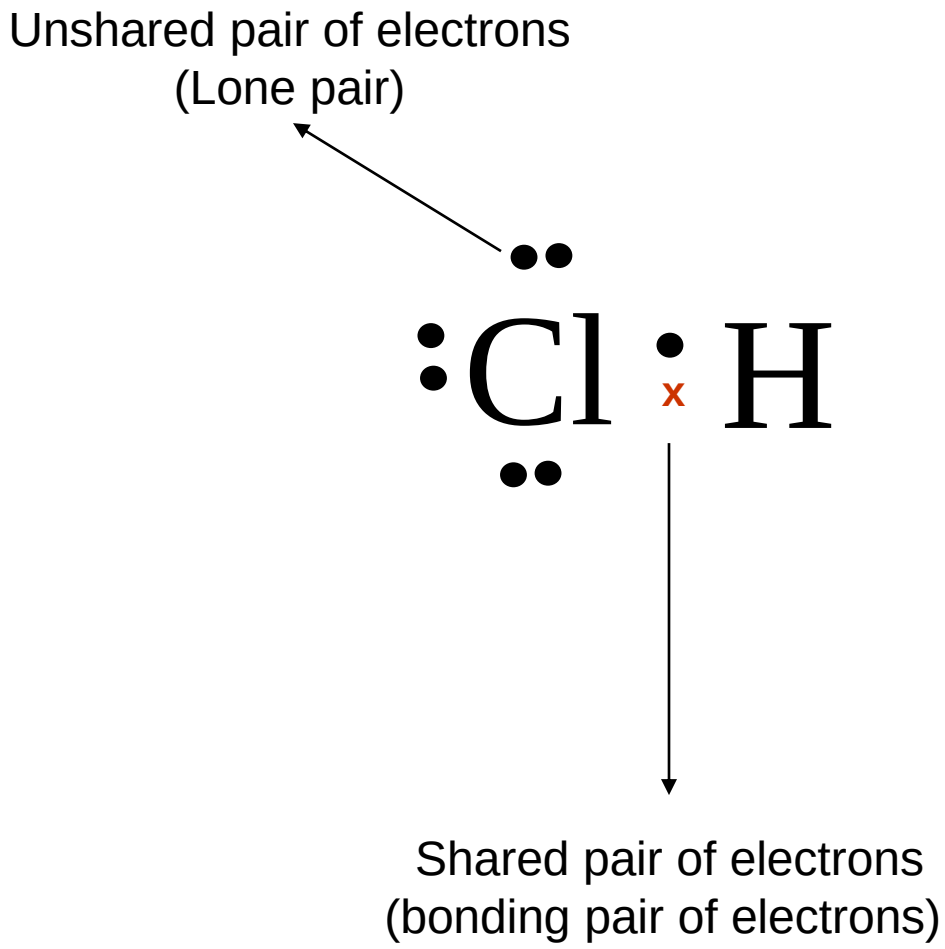


Lewis Dot Structure

Lewis Structure

Cl: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$ ⟶ Ar: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$

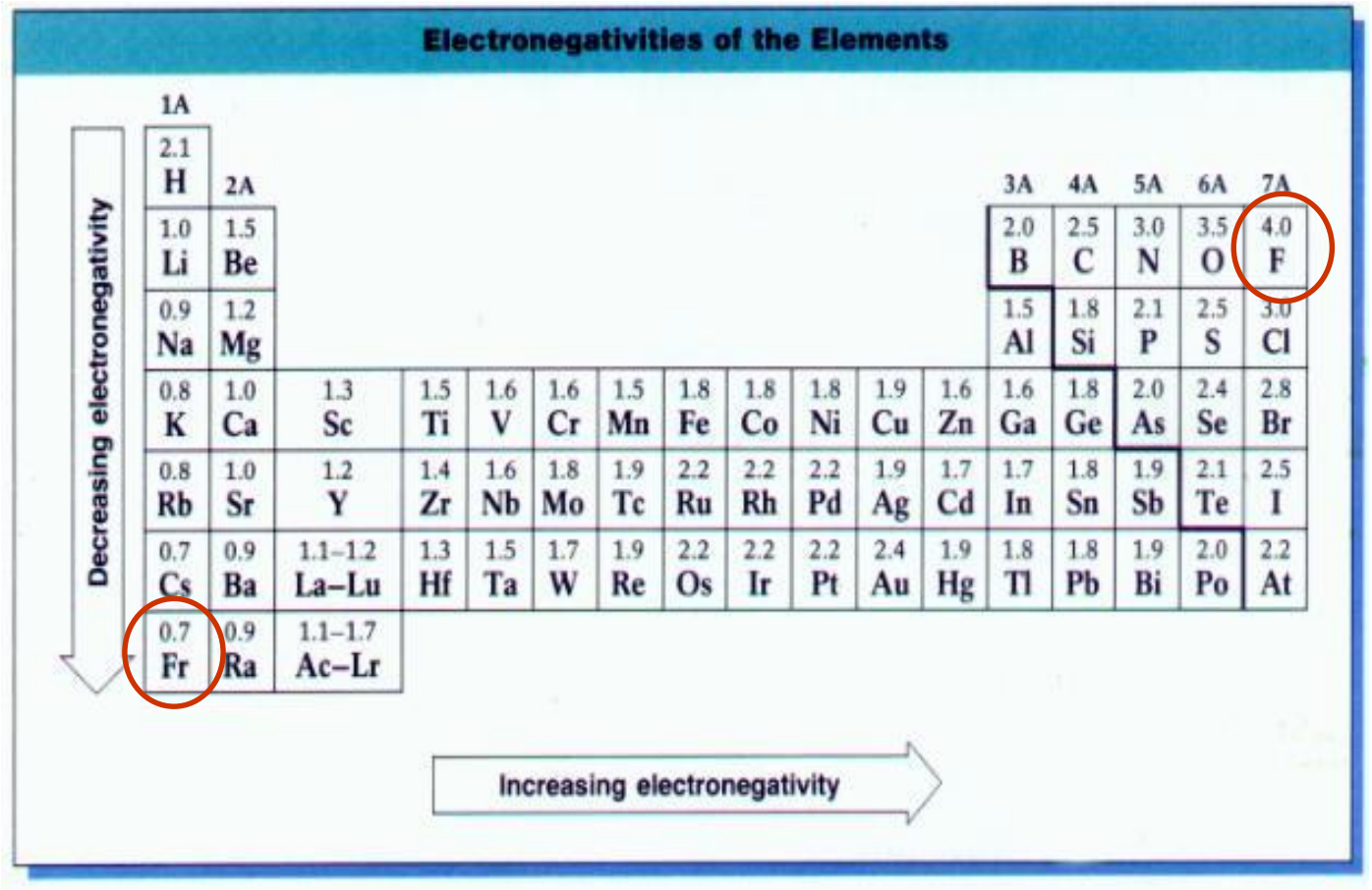
H: $1s^1$ ⟶ He: $1s^2$



Only valance electrons are involved in bonding (ionic and covalent bonds).

Electronegativity

A measure of an atom's attraction for shared electrons.

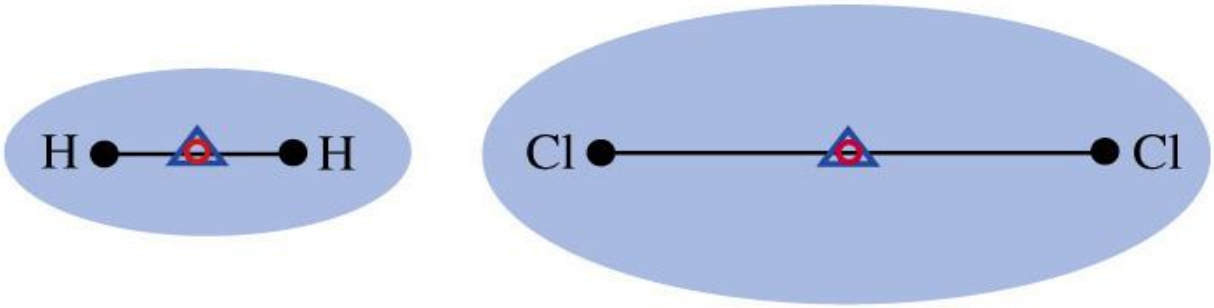


1A	2A	3A	4A	5A	6A	7A
2.1 H						
1.0 Li	1.5 Be	2.0 B	2.5 C	3.0 N	3.5 O	4.0 F
0.9 Na	1.2 Mg	1.5 Al	1.8 Si	2.1 P	2.5 S	3.0 Cl
0.8 K	1.0 Ca	1.3 Sc	1.5 Ti	1.6 V	1.6 Cr	1.5 Mn	1.8 Fe	1.8 Co	1.8 Ni	1.9 Cu	1.6 Zn	1.6 Ga	1.8 Ge	2.0 As	2.4 Se	2.8 Br
0.8 Rb	1.0 Sr	1.2 Y	1.4 Zr	1.6 Nb	1.8 Mo	1.9 Tc	2.2 Ru	2.2 Rh	2.2 Pd	1.9 Ag	1.7 Cd	1.7 In	1.8 Sn	1.9 Sb	2.1 Te	2.5 I
0.7 Cs	0.9 Ba	1.1–1.2 La–Lu	1.3 Hf	1.5 Ta	1.7 W	1.9 Re	2.2 Os	2.2 Ir	2.2 Pt	2.4 Au	1.9 Hg	1.8 Tl	1.8 Pb	1.9 Bi	2.0 Po	2.2 At
0.7 Fr	0.9 Ra	1.1–1.7 Ac–Lr														

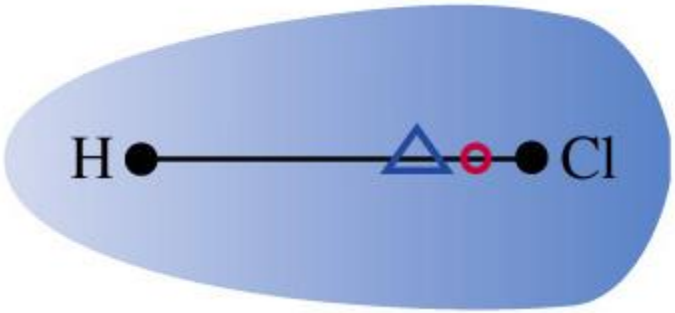
Electronegativity ↑ ⟹ Ionization energy ↑

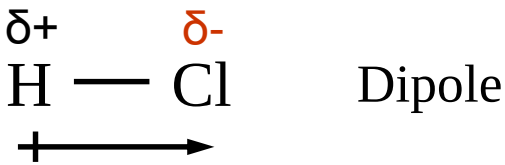
Covalent bonds

Nonpolar covalent bond: electrons are shared equally.

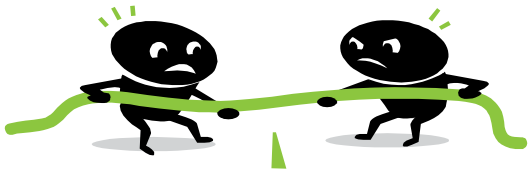


Polar covalent bond: electrons are shared unequally.





Dipole



Electronegativity & bonds

Electronegativity Difference Between Bonded Atoms	Type of Bond
Less than 0.5	Nonpolar Covalent
0.5 to 1.9	Polar Covalent
Greater than 1.9	Ionic

H — H	2.1 – 2.1 = 0	Nonpolar covalent
N — H	3.0 – 2.1 = 0.9	polar covalent
Na — F	4.0 – 0.9 = 3.1	Ionic

Rules to Write Dot Structures

1. Write a skeleton molecule with the lone atom in the middle (Hydrogen can never be in the middle).

2. Find the number of electrons needed (N)
   (8 x number of atoms, 2 x number of H atoms)

1. Find the number of electrons you have (valence e^-'s) (H)

2. Subtract to find the number of bonding electrons (N-H=B)

3. Subtract again to find the number of non-bonding electrons (H-B=NB)

4. Insert minimum number of bonding electrons in the skeleton between atoms only. Add more bonding if needed until you have B bonding electrons.

5. Insert needed non-bonding electrons around (not between) atoms so that all atoms have 8 electrons around them. The total should be the same as NB in 5 above.

Water H_2O

1. S H O H

2. N 2 × 2 = 4 for Hydrogen
1 × 8 = 8 for Oxygen
4+8 = 12 needed electrons

12 N

3. H 2 × 1 = 2 for Hydrogen
1 × 6 = 6 for Oxygen
You have 8 available electrons

- 8 H

4. B 12 - 8 = 4 bonding electrons

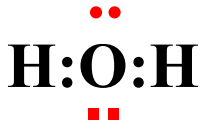
- 4 B

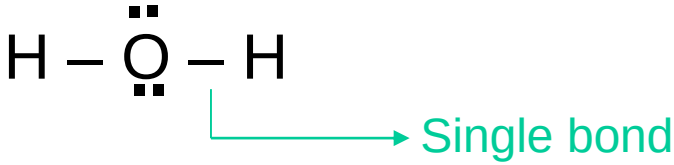
5. NB 8 - 4 = 4 non-bonding electrons

4 NB

6.

H:O:H





Ammonia NH_3

1. S

H
H N H

2. N

3 × 2 = 6 for Hydrogen
1 × 8 = 8 for Nitrogen
6+8 = 14 needed electrons

14 N

3. H

3 × 1 = 3 for Hydrogen
1 × 5 = 5 for Nitrogen
You have 8 available electrons

- 8 H

4. B

14 - 8 = 6 bonding electrons

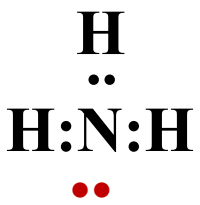
- 6 B

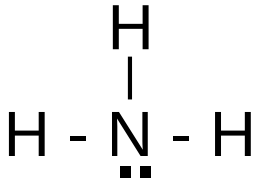
5. NB

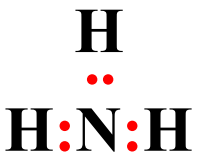
8 - 6 = 2 non-bonding electrons

2 NB

6.







Carbon dioxide CO_2

1. S — O C O

2. N — 1 x 8 = 8 for Carbon
2 x 8 = 16 for Oxygen
8+16=24 needed electrons

24 N

3. H — 1 x 4 = 4 for Carbon
2 x 6 = 12 for Oxygen
You have 16 available electrons

- 16 H

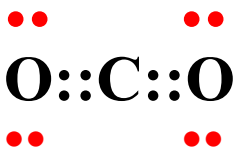
4. B — 24 - 16 = 8 bonding electrons

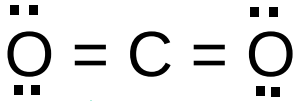
- 8 B

5. NB — 16 – 8 = 8 non-bonding electrons

8 NB

6.





Double bond

O::C::O

Acetylene C_2H_2

1. S H C C H

2. N 2 × 8 = 16 for Carbon
2 × 2 = 4 for Hydrogen
16+4 = 20 needed electrons

20 N

3. H 2 × 4 = 8 for Carbon
2 × 1 = 2 for Hydrogen
You have 10 available electrons

- 10 H

4. B 20 - 10 = 10 bonding electrons

- 10 B

5. NB 10 - 10 = 0 non-bonding electrons

0 NB

6.

Triple bond

Practice

Write the Lewis structure for the:

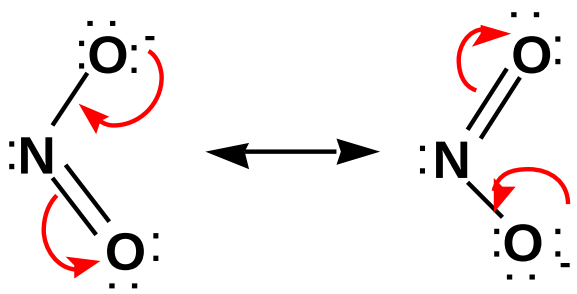
NH_4^+

CN^-

Resonance

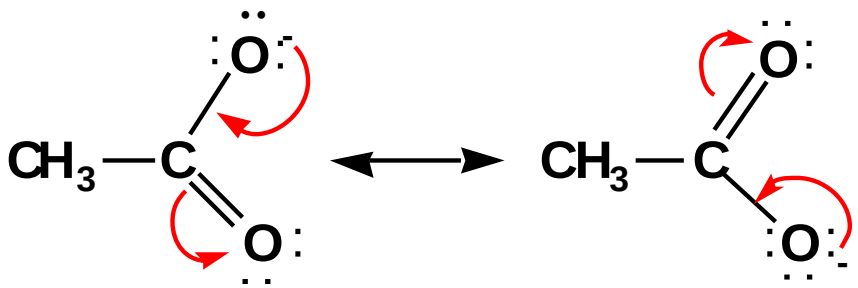
When a molecule has more than one Lewis structure.

NO_2^-

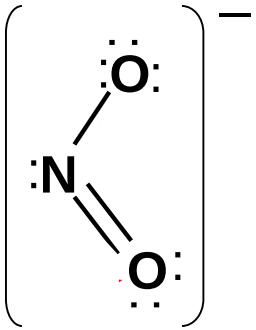


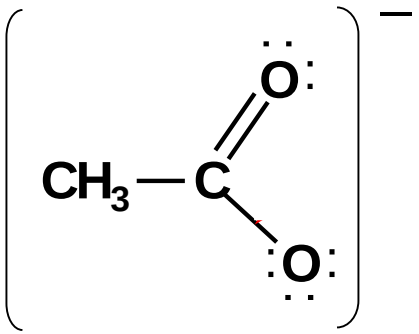
Resonance structures

CH_3COO^-



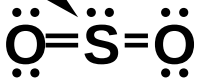
Resonance structures



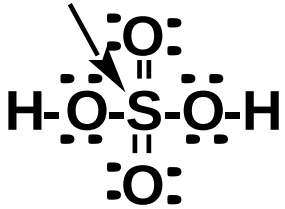


Some exceptions to the Octet rule

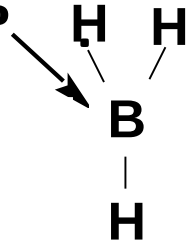
10 electrons in the valence shell of sulfur



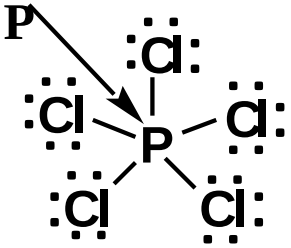
12 electrons in the valence shell of sulfur



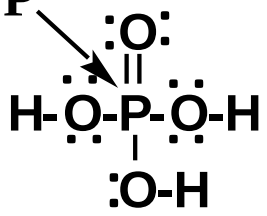
6 electrons in the valence shell of P



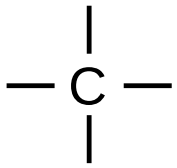
10 electrons in the valence shell of P



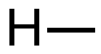
10 electrons in the valence shell of P



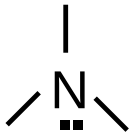
– Carbon: normally forms four covalent bonds and has no unshared pairs of electrons.



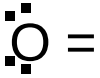
– Hydrogen: forms one covalent bond and no unshared pairs of electrons.



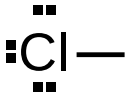
– Nitrogen: normally forms three covalent bonds and has one unshared pair of electrons.



– Oxygen: normally forms two covalent bonds and has two unshared pairs of electrons.



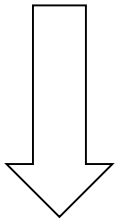
– A Halogen: normally forms one covalent bond and has three unshared pairs of electrons.



VSEPR Model

VSEPR: Valence-Shell Electron-Pair Repulsion method

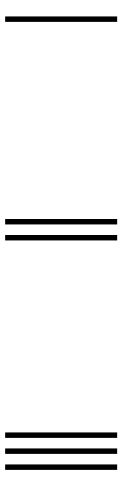
Bond angle: angle between two atoms bonded to a central atom.



Regions of electron like to be
as far away as possible from the others.

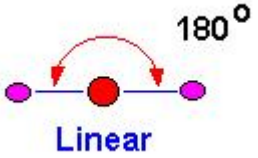
Regions of electron density

Four regions of electron density around an atom:



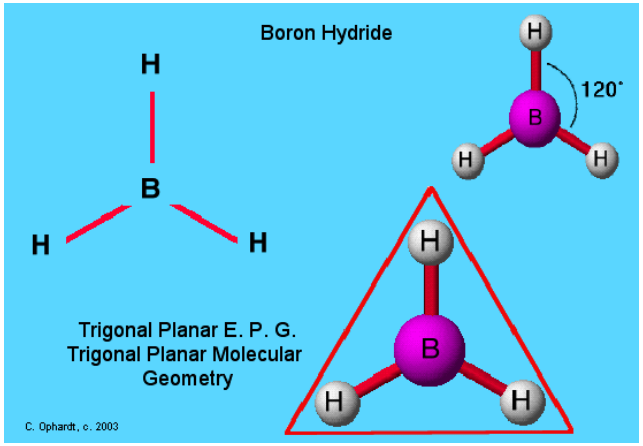
Lone Pair

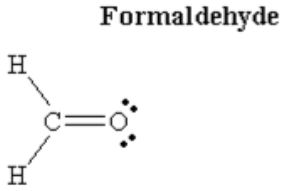
Bond Angles & Geometric Structures

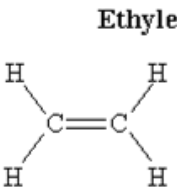


Linear molecules

2 regions

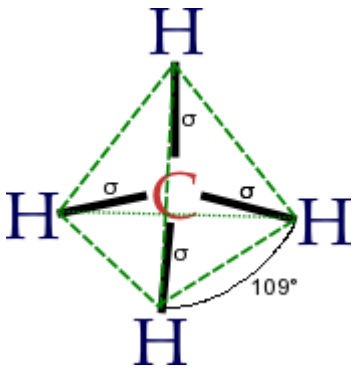


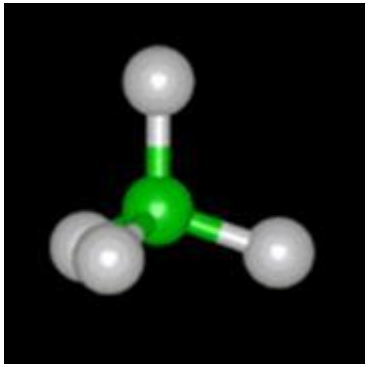




Trigonal planar molecules

3 regions

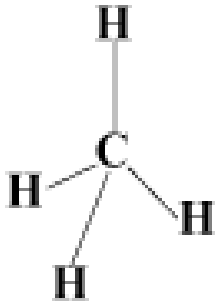




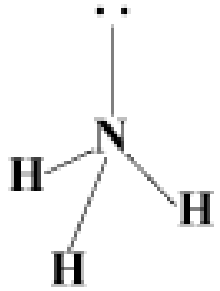
Tetrahedral arrangement

4 regions

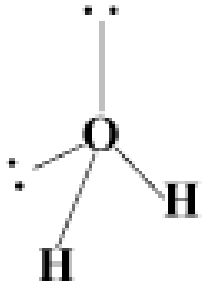
Tetrahedral Electron Pair Geometry
(Molecular Geometry-Shape)



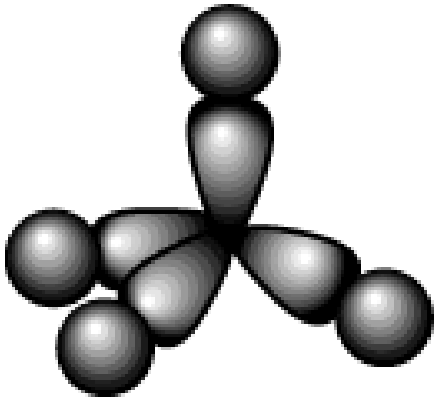
Methane (CH_4)



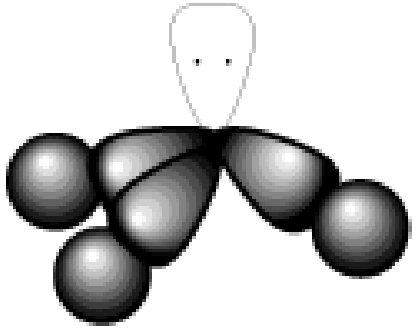
Ammonia (NH_3)



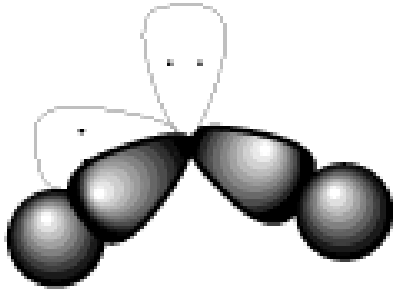
Water (H_2O)



Tetrahedral

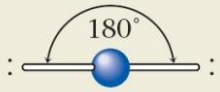
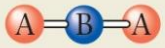
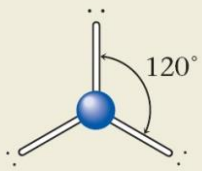
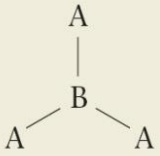
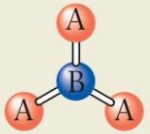
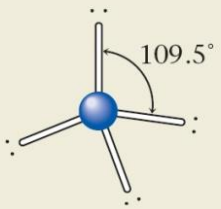
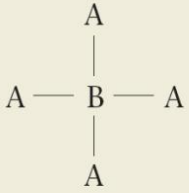
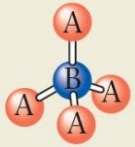
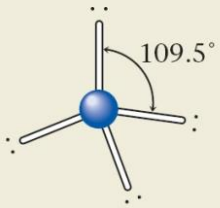
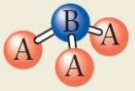
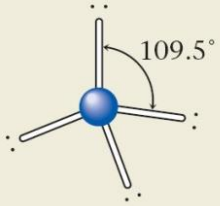
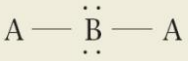
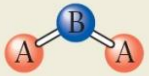


Trigonal Pyramidal



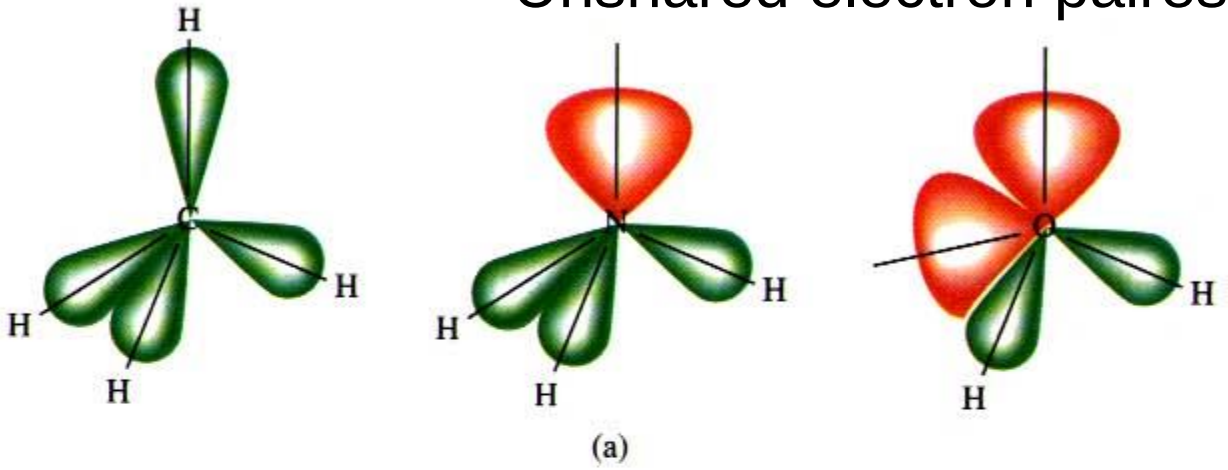
Bent (V-Shaped)

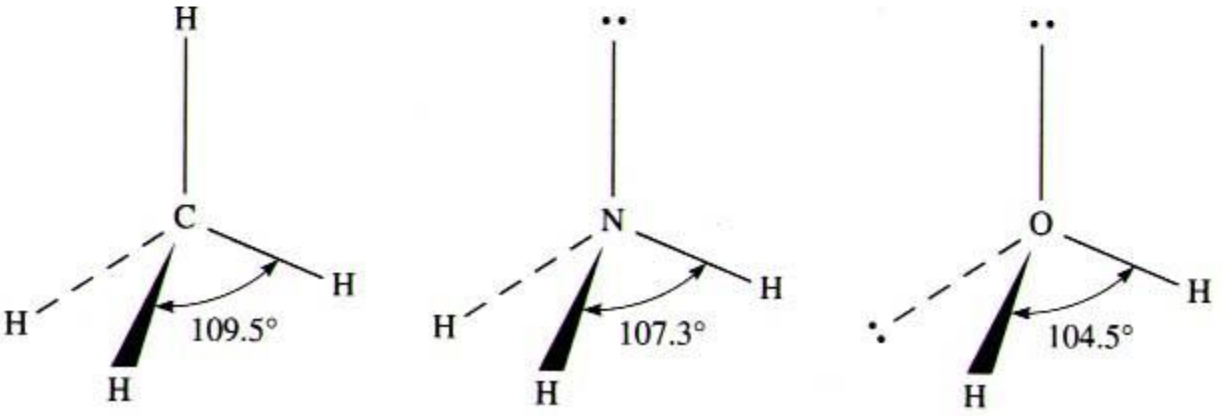
Table 12.4 Arrangements of Electron Pairs and the Resulting Molecular Structures for Two, Three, and Four Electron Pairs

Case	Number of Electron Pairs	Bonds	Electron Geometry (Arrangement)	Ball-and-Stick Model	Angle Between Pairs	Molecular Geometry (Shape)	Partial Lewis Structure	Ball-and-Stick Model	Example
1	2	2	Linear		180°	Linear	A—B—A		BeF_2
2	3	3	Trigonal planar (triangular)		120°	Trigonal planar (triangular)			BF_3
3	4	4	Tetrahedral		109.5°	Tetrahedral			CH_4
4	4	3	Tetrahedral		109.5°	Trigonal pyramid			NH_3
5	4	2	Tetrahedral		109.5°	Bent or V-shaped			H_2O

Tetrahedral Electron Pair Geometry

Unshared electron paires

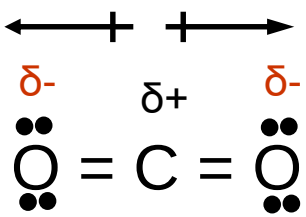




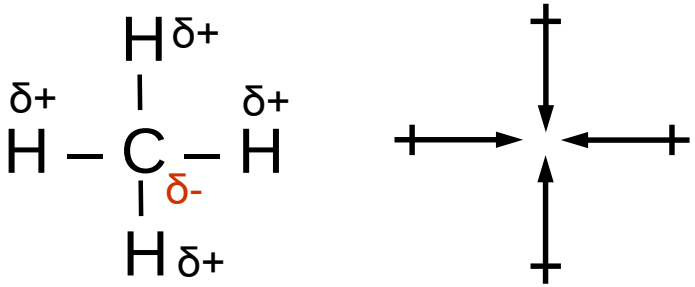
CH_4 NH_3 H_2O

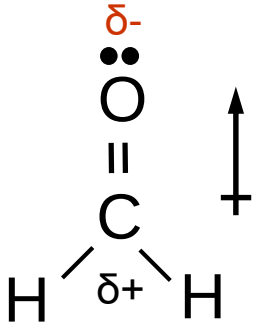
Polarity

1. It has polar bonds.
2. Centers of δ+ and δ- lie at different places (sides).

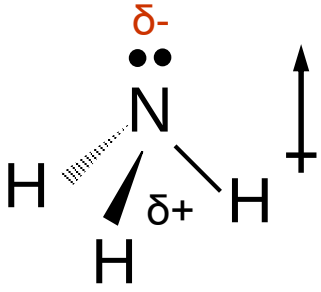


nonpolar molecule





polar molecule



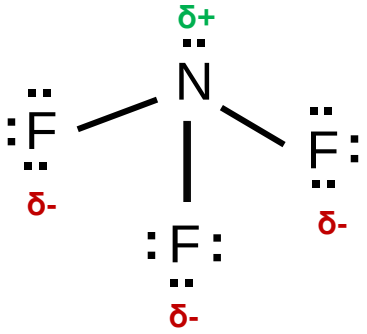
Polarity & shape

Practice:

Use VSEPR theory to predict the molecular structure of these molecules. Which molecule(s) is(are) polar?

NF_3

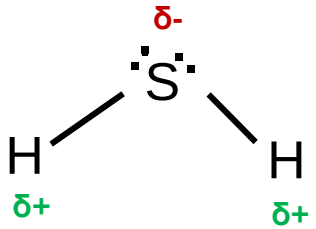
trigonal pyramid



Polar molecule

H_2S

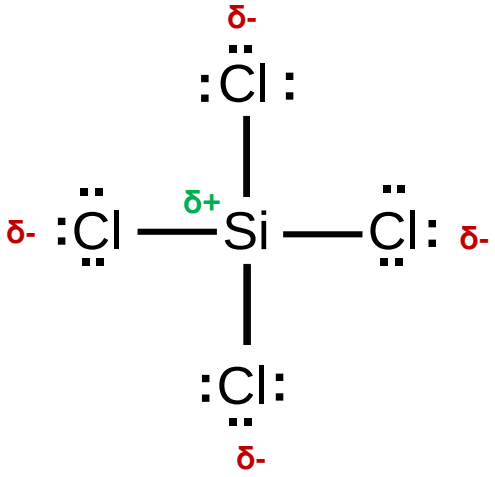
bent



Polar molecule

$SiCl_4$

tetrahedral



Non-polar molecule

At-home Practice

- Predict the molecular shape of **HCN** using the VSEPR model.

- Draw the Lewis structure of **PCl_3**
  - How many lone pairs are on the P atom?
  - Are the phosphorous-chlorine bonds single, double or triple? Are the bonds ionic, polar covalent, or covalent?
  - What is the molecular shape predicted by VSEPR?
  - Based on the molecular shape, is the PCl_3 molecule polar?