


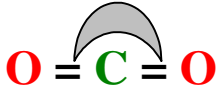
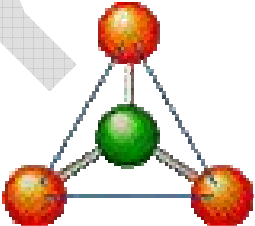
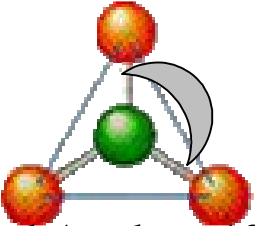
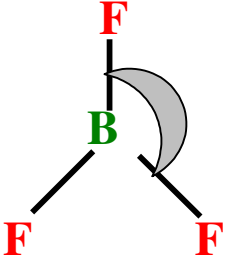
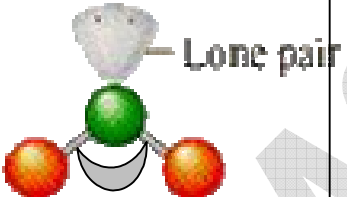

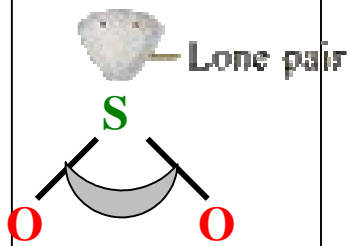
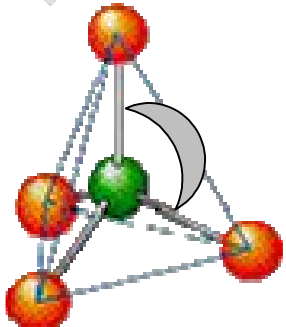
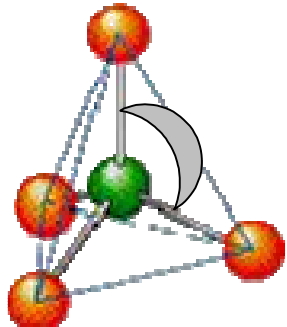
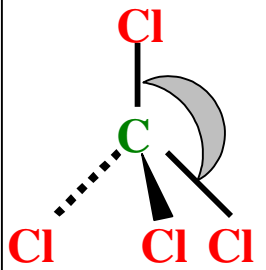
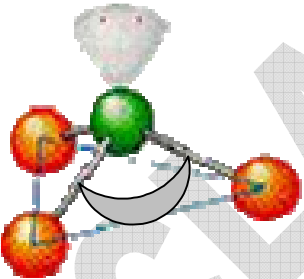
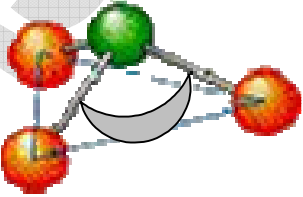

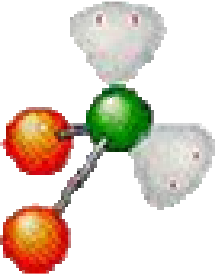
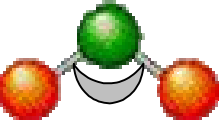
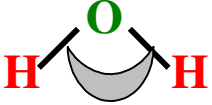
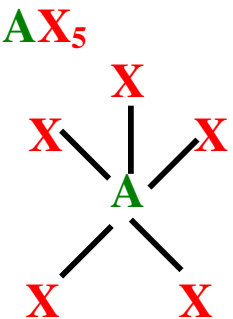
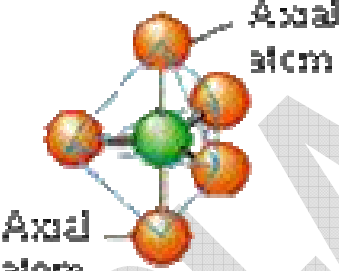
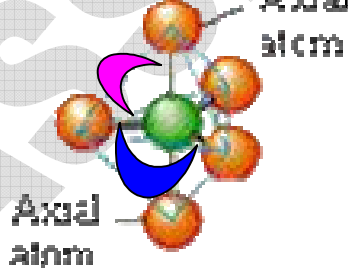
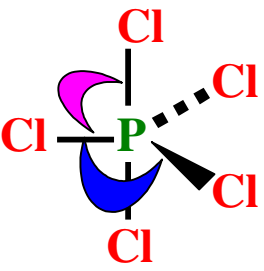
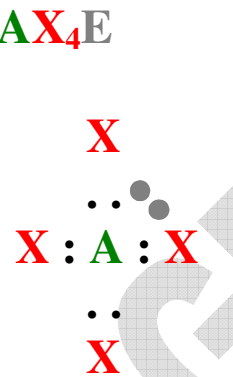
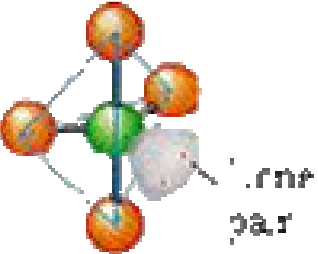
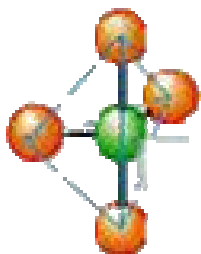
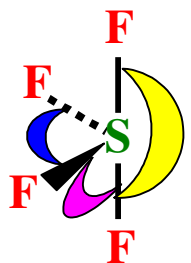


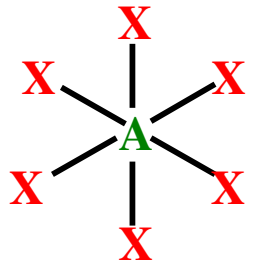
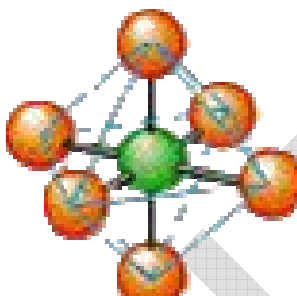
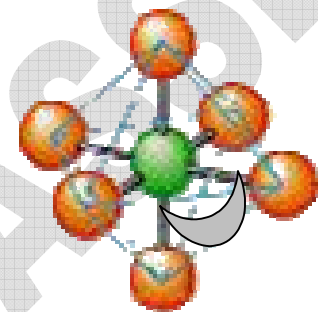
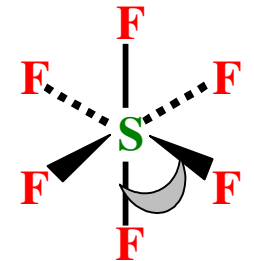
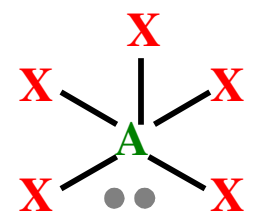
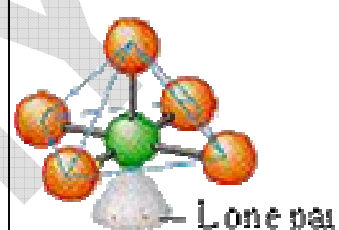
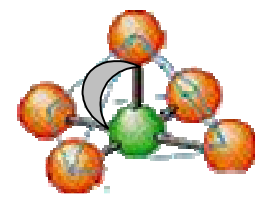
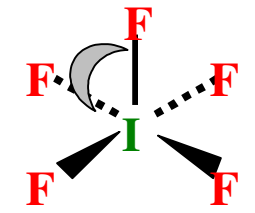
Lewis formula	Electron Pairs			Electron Pair Geometry	Molecular Geometry	Example
	Total	Bonding	Lone			
AX_2 $\text{X} : \text{A} : \text{X}$	2	2	0	 Linear	 Linear Bond Angle = 180° Nonpolar	  Treat double bonds as single bonds
AX_3 X \dots $\text{A} : \text{X}$ \dots X	3	3	0	 Trigonal planar	 Bond Angle = 120° Nonpolar Trigonal Planar	

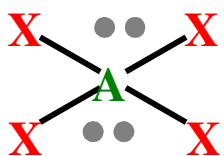
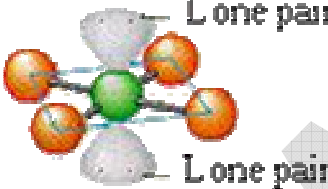
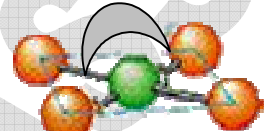
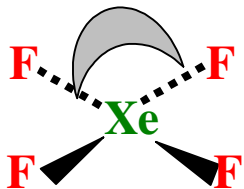
Lewis formula	Electron Pairs			Electron Pair Geometry	Molecular Geometry	Example
	Total	Bonding	Lone			
AX_2E E = lone pair $\begin{array}{c} \cdot\cdot \\ \text{X} : \text{A} :: \text{X} \\ \text{or} \\ \text{X} :: \text{A} : \text{X} \end{array}$ Treat double bonds as single bonds	3	2	1	 Trigonal planar	 Bent (angular) Bond Angle $\approx 120^\circ$ Polar	
AX_4 $\begin{array}{c} \text{X} \\ \cdot\cdot \\ \text{X} : \text{A} : \text{X} \\ \cdot\cdot \\ \text{X} \end{array}$	4	4	0	 Tetrahedral	 Tetrahedral Bond Angle $\approx 109.5^\circ$ Nonpolar	

Lewis formula	Electron Pairs			Electron Pair Geometry	Molecular Geometry	Example
	Total	Bonding	Lone			
AX_3E E = lone pair $\begin{array}{c} \text{X} : \overset{\cdot\cdot}{\underset{\cdot\cdot}{\text{A}}} : \text{X} \\ \text{X} \end{array}$	4	3	1	 Tetrahedral	 Trigonal Pyramidal Bond Angle $\approx 109.5^\circ$ Polar	
AX_2E_2 $\begin{array}{c} \text{X} : \overset{\cdot\cdot}{\underset{\cdot\cdot}{\text{A}}} : \\ \text{X} \end{array}$	4	2	2	 Tetrahedral	 Bent (Angular) Bond Angle $\approx 109.5^\circ$ Polar	


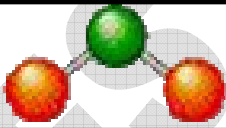
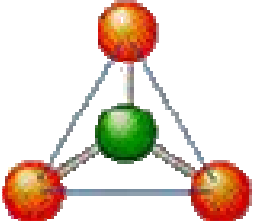
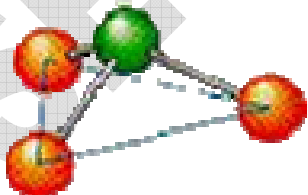
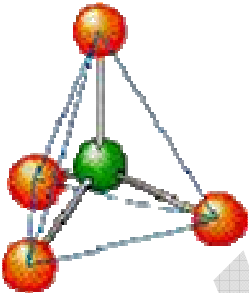
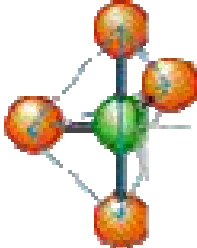
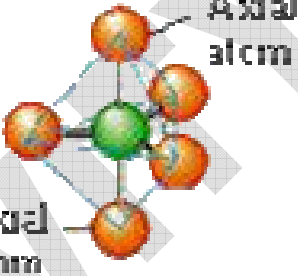
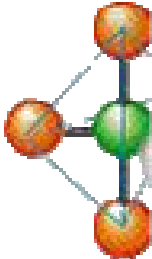
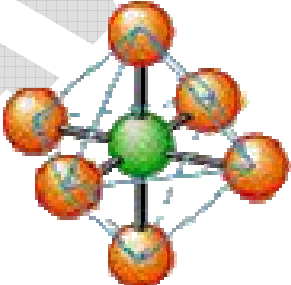
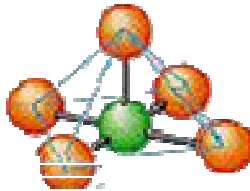

Lewis formula	Electron Pairs			Electron Pair Geometry	Molecular Geometry	Example
	Total	Bonding	Lone			
AX_5 	5	5	0	 Trigonal bipyramidal	 Trigonal bipyramidal Bond Angles: 120° (3) and 90° (6) Nonpolar	 PCl_5
AX_4E 	5	4	1	 Trigonal bipyramidal	 Seesaw Bond Angles: 120° , 90° , and 180° Polar	SF_4 

Lewis formula	Electron Pairs			Electron Pair Geometry	Molecular Geometry	Example
	Total	Bonding	Lone			
AX_3E_2 	5	3	2	<p>Trigonal bipyramidal</p>	<p>T-shaped Bond Angles: 90° Polar</p>	ClF_3
AX_2E_3 	5	2	3	<p>Trigonal bipyramidal</p>	<p>Linear Bond Angle = 180° Nonpolar</p>	XeF_2

Lewis formula	Electron Pairs		Electron Pair Geometry	Molecular Geometry	Example	
	Total	Bonding Lone				
AX_6 	6	6	0	 Octahedral	 Octahedral Bond Angle = 90° Nonpolar	SF_6 
AX_5E 	6	5	1	 Octahedral	 Square pyramidal Bond Angle: 90° Polar	IF_5 

Lewis formula	Electron Pairs		Electron Pair Geometry	Molecular Geometry	Example	
	Total	Bonding Lone				
AX_4E_2 	6	4	2	 Octahedral	 Square planar Bond Angle: 90° Nonpolar	XeF_4 

SUMMARY

Symmetrical Arrangement	Asymmetrical Arrangement
Polar Bonds \Rightarrow Nonpolar Molecule	Polar Bonds \Rightarrow Polar Molecule
Linear 	Bent 
Trigonal Planar planar 	Trigonal pyramidal 
Tetrahedral 	Seesaw 
Trigonal bipyramidal 	T-shaped 
Octahedral 	Square pyramidal 
Square planar 	

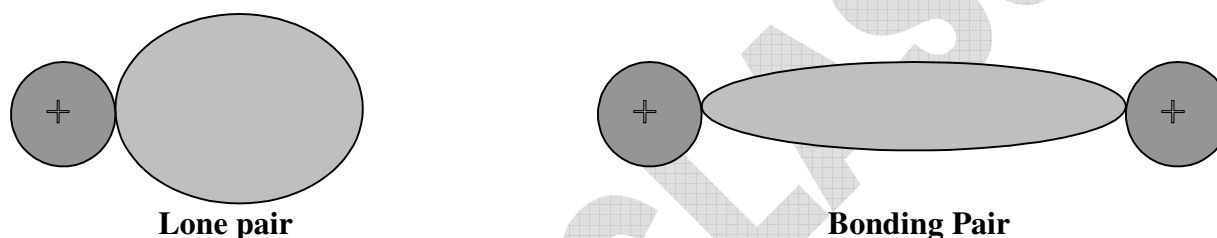
BOND ANGLES

- Bond angles between atoms can be approximately predicted from the VSEPR model.
- Some deviations from the predicted bond angles have been determined experimentally. These deviations are caused by 2 factors:

1. Effect of Lone Pairs

Alone pair tends to require more space than a bonding pair

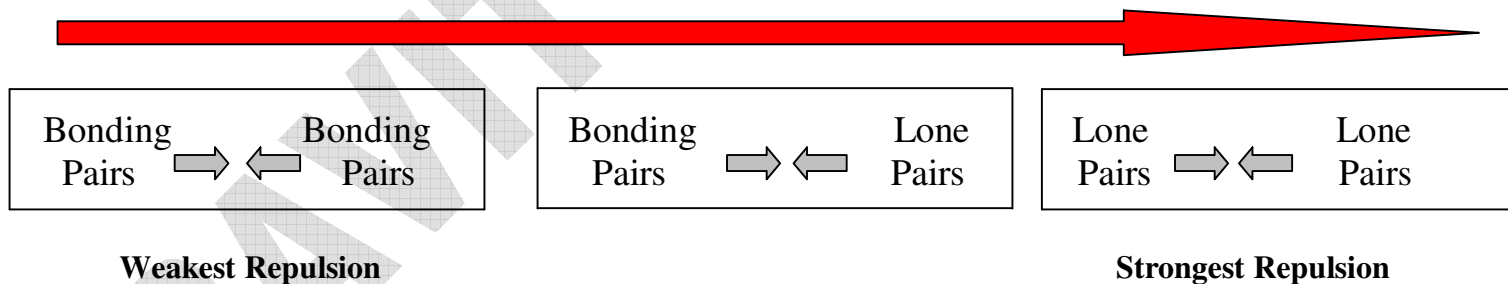
Reason: A lone pair of electrons is attracted to only one atomic core, whereas a bonding pair is attracted to two.
The lone pair is larger, while the bonding pair is drawn more tightly to the nuclei.

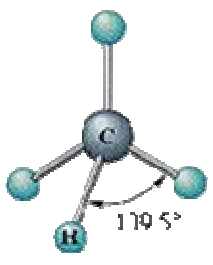
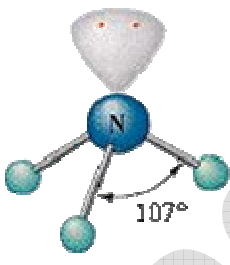
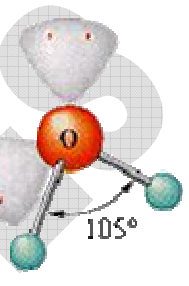


Lone pairs repel each other stronger than Bonding Pairs.

Result: The repulsions between electron pairs depend on the type of electron pairs involved.

REPULSION INCREASES



	CH ₄	NH ₃	H ₂ O
			
Electron Pair Geometry	Tetrahedral	Tetrahedral	Tetrahedral
Expected Bond Angle	109.5°	109.5°	109.5°
Bonding Pairs	4	3	2
Lone Pairs	0	1	2
Actual Bond Angle	109.5°	107°	105°

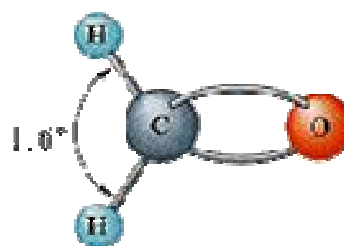
2. Effect of Multiple Bonds

Multiple bonds require more space than single bonds because of the greater number of electrons:

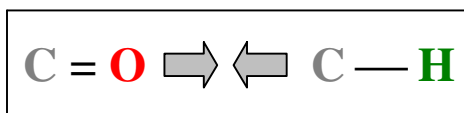
Electron Pair Geometry: Trigonal Planar

Expected Bond Angle: 120°

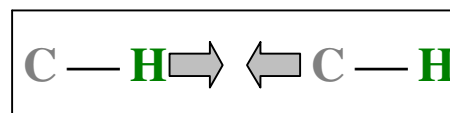
Actual Bond Angle: 116°



Reason:



Stronger Repulsion



Weaker Repulsion