Chemistry 11

Atomic Theory IV

Name: Key
Date:
Block:

> 1. **Lewis Diagrams**
> 2. **VSEPR**

Lewis Diagrams

- Lewis diagrams show the bonding between atoms of a molecule.
- Only the outermost electrons of an atom (called Valence electrons) are involved in bonding (usually just p and s)

Fill in the chart below to determine the valence electrons of elements 3-10

Element	Lithium	Beryllium	Boron	Carbon	Nitrogen	Oxygen	Fluorine	Neon
Group #	1	2	13	14	15	16	17	18
Full Electron Configuration	$1s^2 2s^1$	$1s^2 2s^2$	$1s^2 2s^2 2p^1$	$1s^2 2s^2 2p^2$	$1s^2 2s^2 2p^3$	$1s^2 2s^2 2p^4$	$1s^2 2s^2 2p^5$	$1s^2 2s^2 2p^6$
# of Valence Electrons	1	2	3	4	5	6	7	8

In general:

Main Group Number	1	2	13	14	15	16	17	18
Valence Electrons	1	2	3	4	5	6	7	8
Valence Electron Configuration	ns^1	ns^2	ns^2np^1	ns^2np^2	ns^2np^3	ns^2np^4	ns^2np^5	ns^2np^6

* n = energy level

When drawing Lewis dot structures:

- Draw 1 dot for each valence electron
- Begin pairing dots only after you have put a dot on each side (north, east, south, west) of the atom

Draw the Lewis structures for the elements belonging to <u>period 4</u> of the periodic table:

Group 1	Group 2	Group 13	Group 14	Group 15	Group 16	Group 17	Group 18
K	Ca	Ga	Ge	As	Se	Br	Kr

Draw the Lewis structures for the following atoms and ions:

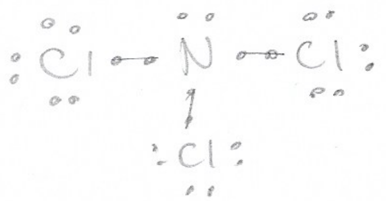
Ba	Br	Br^-	Bi	Al^{3+}	Te
Ba	Br	$[Br]^-$	Bi	$[Al]^{3+}$	Te

Draw the Lewis structures for the ions of these elements:

Ca	Se	Ga	As	Cl
$\left[:\ddot{\underset{..}{Ca}}:\right]^{2+}$	$\left[:\ddot{\underset{..}{Se}}:\right]^{2-}$	$\left[:\ddot{\underset{..}{Ga}}:\right]^{3+}$	$\left[:\ddot{\underset{..}{As}}:\right]^{3-}$	$\left[:\ddot{\underset{..}{Cl}}:\right]^{-}$

Lewis Structures for Molecules:

Example: NCl_3

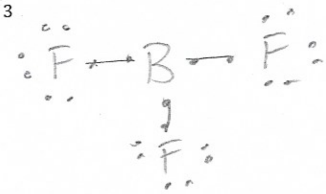
What to Think About	How to Do It
1. Figure out the total number of valence electrons in the molecule	N=5 Cl=7x3=21 5+21=$\boxed{26}$
2. Arrange the atoms. Assume that hydrogen and the halogens will not be the central atom	Cl N Cl Cl
3. Draw valence electrons around each atom	
4. Connect unpaired electrons with a bond. Remember: there are two electrons in every bond. Some molecules may need double bonds. - *H atoms form only one bond* - *O normally forms two bonds* - *N normally forms three bonds* - *C normally forms four bonds* - *Halogens normally form only one bond*	
5. Redraw the diagram from step 4 neatly.	
6. Do a final check: ☐ **Do all the valence electrons in the diagram (bonds AND dots) match the total number of valence electrons from step 1?** ☐ **Do all atoms follow the octet rule (8 electrons in the valence shell)?**	✓ ✓

Exceptions to the OCTET RULE:

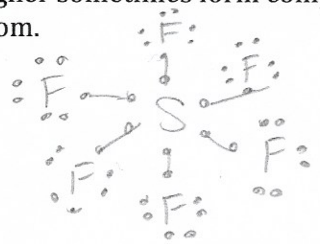
1. The incomplete octet
   - Elements in groups 1, 2 and 13 tend to form compounds in which they are surrounded by fewer than eight electrons
   - Examples:

H_2O

BF_3

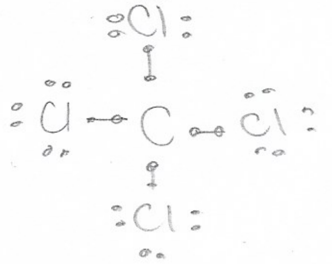
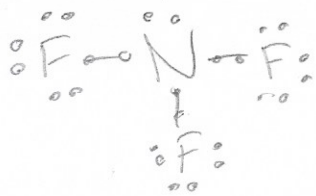


2. The expanded octet
   - Atoms in period 3 or higher sometimes form compounds in which more than eight electrons surround the central atom.
   - Example: SF_6



3. Odd-electron molecules
   - Some molecules contain an odd number of electrons.
   - Odd-electron molecules are called radicals
   - Example: NO_2

With the above steps, construct Lewis structures for the following:

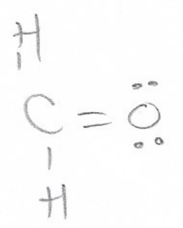
1. CCl_4 	2. NF_3 
3. H_2O 	4. H_2Se 
5. NH_3 	6. OF_2 

Lewis Structures for Molecules with Multiple Bonds:

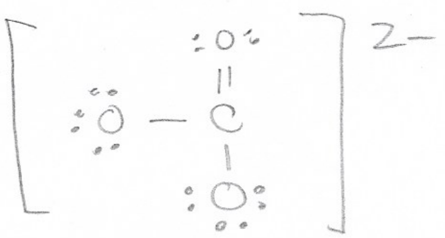
1. CO_2

2. SO_2

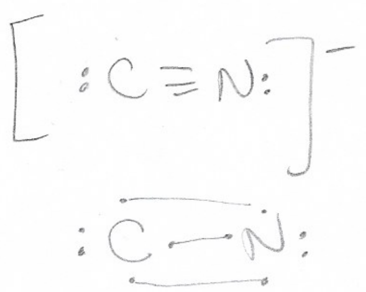
3. CH_2O (*hint: Carbon is the central atom)



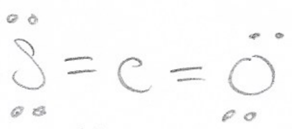
4. CO_3^{2-} (*hint: **2-** adds 2 electrons to the total number of valence electrons)



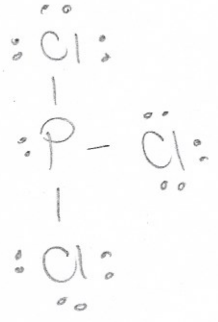
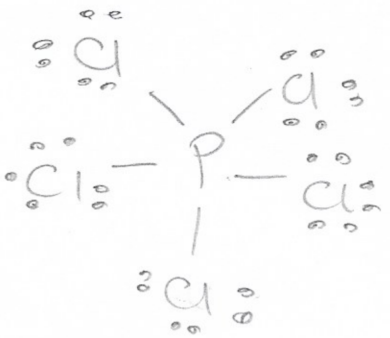
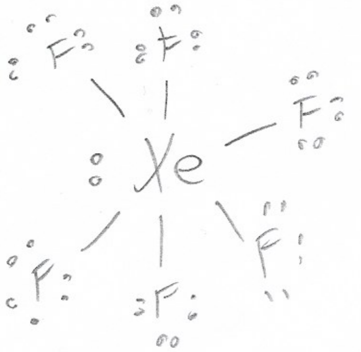
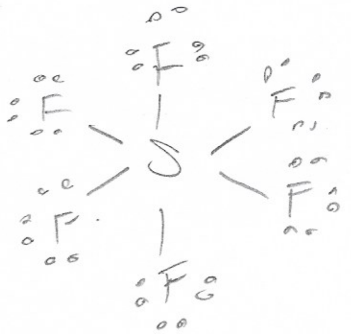
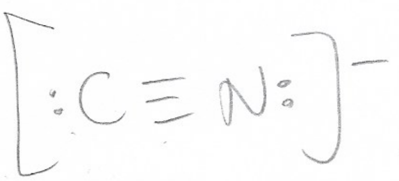
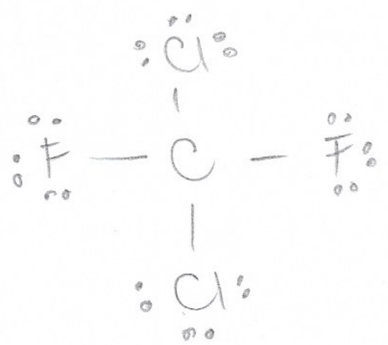
5. CN^-(*hint: **-** adds 1 electron to the total number of valence electrons)



6. SCO (*hint: C is the central atom)



More Practice!!

1. PCl_3 	2. PCl_5 
3. XeF_6 	4. SF_6 
5. CN^- 	6. CCl_2F_2 

VSEPR (Valence Shell Electron Pair Repulsion)

In order to understand the shapes the molecules form we must adhere to the same rules we have been following throughout this section

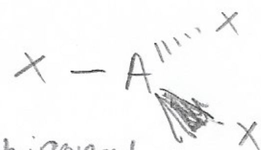
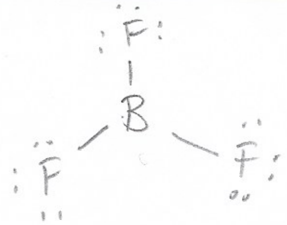
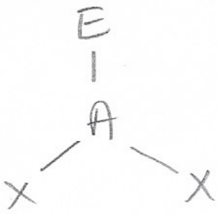
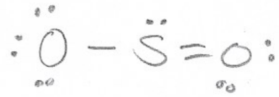
i. Electrons all have the same negative charge
ii. Like charges repel
iii. Bonded pairs surrounding the nucleus repel other bonded pairs and other electrons
iv. Lone pairs surrounding the nucleus repel other bonded pairs and other electrons
v. Valence electrons are oriented in such a way as to be as far from each other as possible

A = central atom **X =** ligands (atoms bonded to A) **E =** electron pair

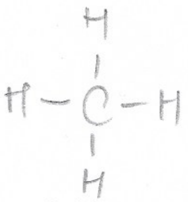
Two-Bonding Electron Groups: AX_2

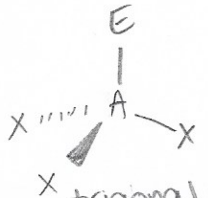
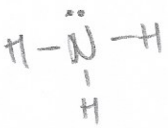
Notation	Molecular Shape	Sample Lewis Structure
AX_2	X — A — X Linear	 carbon dioxide

Three-Electron Groups: AX_3 and AX_2E

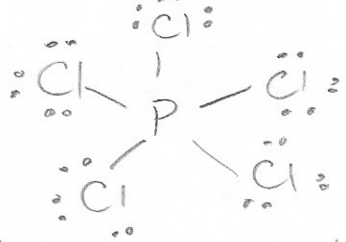
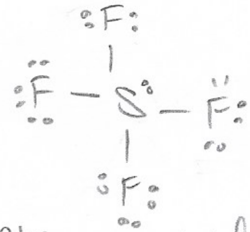
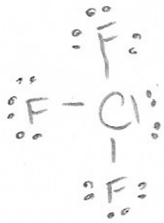
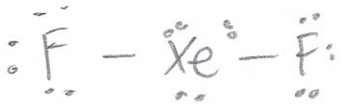
Notation	Molecular Shape	Sample Lewis Structure
AX_3	 trigonal planar	
AX_2E	 bent	 Sulfur dioxide

Four-Electron Groups: AX_4, AX_3E and AX_2E_2

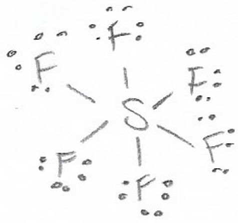
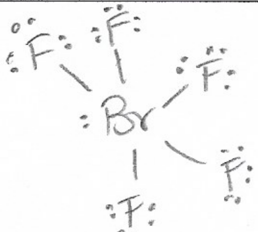
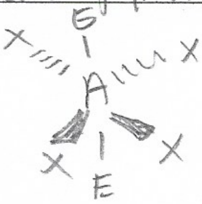
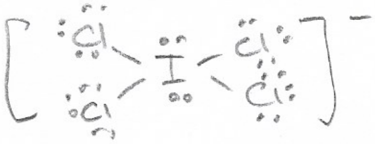
Notation	Molecular Shape	Sample Lewis Structure
AX_4	 Tetrahedral	 methane

AX_3E	 trigonal pyramidal	 ammonia
AX_2E_2	 bent	 water

Five-Electron Groups: AX_5, AX_4E, AX_3E_2, AX_2E_3

Notation	Molecular Shape	Sample Lewis Structure
AX_5	 triganol bipyramidal	 phosphorus pentachloride
AX_4E	 Seesaw	 Chlorine triflu ride
AX_3E_2	 T-shaped	 Chlorine trifluoride
AX_2E_3	 Linear	 Xenon difluoride

Six-Electron Groups: AX_6, AX_5E, AX_4E_2

Notation	Molecular Shape	Sample Lewis Structure
AX_6	 octahedral	
AX_5E	 Square pyramidal	 bromine pentafluoride
AX_4E_2	 Squareplanar	 tetrachloroiodate ion

VSEPR Nomenclature						
		0 lone pairs	1 lone pair	2 lone pairs	3 lone pairs	4 lone pairs
Number of ligands around the central atom	2	*linear*				
	3	*trigonal planar*	*bent*			
	4	*tetrahedral*al	*trigonal pyramid*al	*bent*		
	5	*trigonal bipyramid*	*seesaw*	*T-shape*	*linear*	
	6	*octahedral*	*square pyramid*al	*square planar*	*T-shape*	*linear*