Unit #1 -Module #6 – Lecture Note – Lewis Dot Diagrams:

Lesson and Guided Learning:

Covalent bonding Lewis Dot Diagrams: (electrons are shared)

A **Lewis dot diagram** is a representation of the valence electrons of an atom that uses **dots (or x's)** around the symbol of the element. The number of **dots (x's)** equals the number of valence electrons in the atom.

Valence Electrons – number of electrons in the outermost energy level (these are involved in bonding)

Table 1: Atoms

Atom	Group	Valence Electrons	Lewis Dot Diagram	# of available Bonding spots	# of lone pairs
N					
O					
F					
H					
B					
O^{2-}					
Br^-					

Terminology:

Bonding spots: Number of places available for bonding to take place.

Lone Pairs: a pair of electrons in the outermost shell of an atom that are not used in bonding

Bonding Pairs: a pair of electrons that is shared to create a bond

Octet: a grouping of eight electrons in an outer energy level of an atom (usually implies stability);

Sub-octet: a group of less than eight electrons in an outer energy level of an atom (stability occurs if hydrogen has two electrons/ boron has six otherwise for the most part unstable)

Single Bonds – two electrons are shared in a bond

Structural diagram – the "shared" electrons are replaced with a line (rather than dots/x's)

Table 2: Bimolecular (two non-metal atoms in a bond)

Molecule	Total available electrons	Lewis Dot Diagram	Structural diagram
Cl_2	2 x 7 e- = 14 e-		
H_2			
F_2			

Table 3: Molecules with a Central Atom

- The central atom is usually the atom with the most bonding sites available;
- The central atom can also be shown by being underlined;

Molecule	Available Electrons	Lewis Dot Diagram	Lone Pairs on the central atom	Bonding pairs on the central atom	Structural Diagram
CF_4					

H_2S					
PH_3					
BCl_3					
H_2O					

Table 4: Molecules with multiple bonds

- **Multiple bond** is sharing 4 or 6 electrons;
- **Double bond** – four electrons are shared in a bond; **Triple bond** – six electrons are shared in a bond;

Molecule	Available Electrons	Lewis Dot Diagram	Lone Pairs on the central atom	Bonding pairs on the central atom	Structural Diagram
CO_2					
HCN					
C_2H_4					

Table 5: Polyatomic Ions

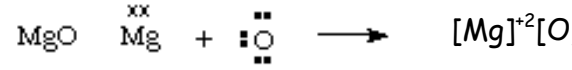
- Molecules with a positive or negative charge

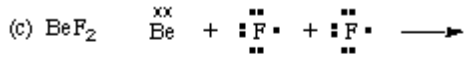
Molecule	Available Electrons	Lewis Dot Diagram	Lone Pairs on the central atom	Bonding pairs on the central atom	Structural Diagram
PH_4^+					

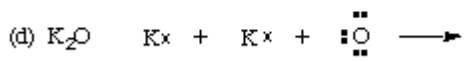
$\underline{Cl}O^-$					

Ionic Bonding Lewis Dot Diagrams: (Transfer of Electrons)

NaF  $[Na]^+[F]^-$

MgO  $[Mg]^{+2}[O]^{-2}$

(c) BeF_2  $[Be]^{+2}[F]_2^-$

(d) K_2O  $[K]_2^+[O]^{-2}$

Polar Bonds:

Electronegativity is the measure of the tendency of an atom to attract a bonding pair of electrons;
Trend in EN – increases up and to the right towards fluorine (being the highest at 4.0)
Electronegativity difference (ΔEN) is the difference in electronegativities of two bonded atoms.
Non-polar covalent bond is when the (ΔEN) is 0
Polar covalent bond is when the (ΔEN) is greater than 0 but less than 1.7 (may differ slightly in biology);
Ionic bond is when the (ΔEN) is greater than 1.7.
Dipole is difference in electronegativity between the different atoms;
Bond dipole difference in electronegativity along a bond.

Keeping it simple:
(Document Camera examples)

Individual Learning

1. Fill in the following chart:

Molecule	Available Electrons	Lewis Dot Diagram	Lone Pairs on the central atom	Bonding pairs on the central atom	Structural Diagram
$CHCl_3$					
N_2					
OCl_2					
NI_3					
OH^-					
OCN^-					
NH_4^+					
C_2F_2					
C_2Br_4					

Ionic Bonds:

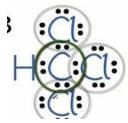
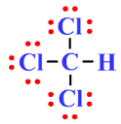
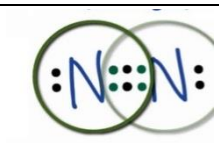
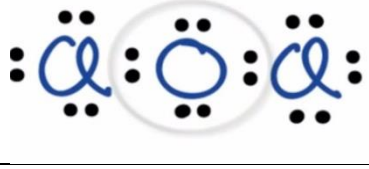
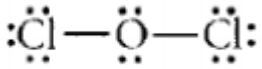
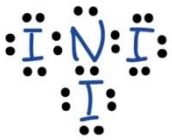
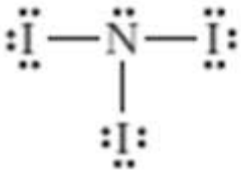
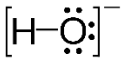
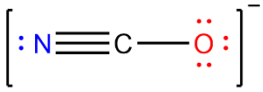
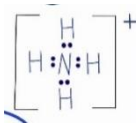
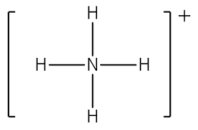
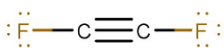
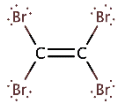
Draw the Lewis Dot Diagrams for the following:

1) KF
2) $MgCl_2$

Polar Bonds:

1. Draw the following bonds and label the bond dipoles for each bond. Indicate whether it is a polar or non-polar bond: a) H-Cl b) C-H c) N-O d) I-Br

ANSWERS to INDIVUAL LEARNING: (No Peeking – until you've tried it!)

Molecule	Available Electrons	Lewis Dot Diagram	Lone Pairs on the central atom	Bonding pairs on the central atom	Structural Diagram
$CHCl_3$	1 x 4e = 4 1 x 1e = 1 3 x 7e = 21 = 26 e		0	4	
N_2	2 x 5e = 10 e		N/A	N/A	
OCl_2	1 x 6e = 6 2 x 7e = 14 = 20e		2	2	
NI_3	1 x 5e = 5 3 x 7e = 21 = 26e		1	3	
OH^-	1 x 6e = 6 1 x 1e = 1 = 7 + 1 = 8e		N/A	N/A	
OCN^-	1 x 6e = 6 1 x 4e = 4 1 x 5e = 5 = 15 + 1 = 16		0	2(triple+single)	
NH_4^+	1 x 5e = 5 4 x 1e = 4 = 9 – 1 = 8		0	4	
C_2F_2	2 x 4e = 8 2 x 7e = 14 = 22e		N/A	N/A	
C_2Br_4	2 x 4e = 8 4 x 7e = 28 = 36e		N/A	N/A	

Ionic Bonds:

Draw the Lewis Dot Diagrams for the following:

1) $KF \rightarrow [K]^{+} [F]^{-}$

2) $MgCl_2 \rightarrow [Mg]^{+2}[Cl]_2^{-}$

Polar Bonds:

1. Draw the following bonds and label the bond dipoles for each bond. Indicate whether it is a polar or non-polar bond: a) H-Cl b) C-H c) N-O d) I-Br

a) H-Cl

$\delta^+ \rightarrow \delta^-$

polar covalent

b) C-H

$\delta^- \leftarrow \delta^+$

polar covalent

c) N-O

$\delta^+ \rightarrow \delta^-$

polar covalent

d) I-Br

$\delta^+ \rightarrow \delta^-$

polar covalent