SCH 3UI Unit 3 Outline:
Chemical Bonding

Lesson	Topics Covered	Homework Questions and Assignments
review	**Note: Electron Configurations, Rutherford-Bohr and Electron Dot Diagrams (EDDs)**	• complete handout and homework: Electron Configurations, Rutherford-Bohr and Electron Dot Diagrams
1	**Note: Chemical Bonding** • review of concepts for chemical bonding, take up EDDs • octet rule • ionic bonding using electron dot diagrams (show 3 steps) • physical properties of ionic compounds • crystal lattice structure	• use EDDs to show the formation of the ionic compounds between: a) Li and P b) Sc and N c) Ba and O d) Al and S • answer questions 9, 12, 13 on pages 73 – 74 in text • define: octet rule, ionic compound, ionic bond, formula unit, crystal lattice (sketch?)
2	**Note: Covalent Bonding** • bonding capacity • diatomic elements (HOBrFINCℓ) • multiple bonds • carbon compounds	• complete homework at the bottom of the handout: Covalent Bonding
3	**The Shapes of Molecules** • AXnEm	• complete Homework: The Shapes of Molecules in the handouts
4	**Note: Electronegativity and Bonding** • electronegativity and chemical bonding (tug of war demo, note) • non-polar covalent bonding • polar covalent bonding • ionic bonding • the bonding continuum	• read and complete practice questions on worksheet: Classifying Bonds and the Bonding Continuum • continue practising drawing Lewis structures to show covalent bonding
5	**Determining the Polarity of Molecules** • consider BOTH the symmetry of the molecule and the polarity of the bonds	• homework at the bottom of the handout: Practice Determining the Polarity of Molecules
6	**Note: Properties of Ionic and Covalent Compounds** • intra- vs inter-molecular forces of attraction **Prelab** for Lab #4	• UNDERSTAND the distinction between intra- and inter-molecular forces of attraction
7	**Lab #4** : Properties of Ionic and Covalent Compounds	• begin lab report for Lab #4.
8	**Note: The Effects of Polarity on the Physical Properties of Compounds** • the bonding continuum • how the properties of compounds are related to the nature of bonding and the polarity of molecules • hydrogen bonding	• continue working on lab report for Lab #4 • complete handout from class: The Effects of Polarity on the Physical Properties of Compounds • For the compounds on the handout "Practice Determining the Polarity of Molecules" (Lesson 05), use the polarity of the molecules to determine the strength of inter-molecular attraction and predict the physical properties for each compound • work on Review for Unit Test #2: Chemical Bonding
9	**Unit Test: Chemical Bonding**	Date:

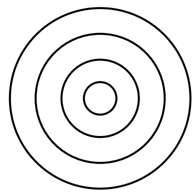
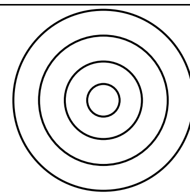
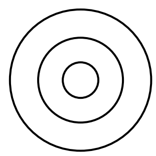
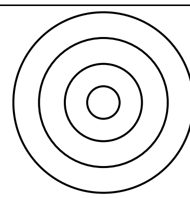
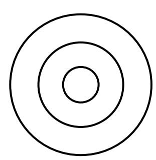
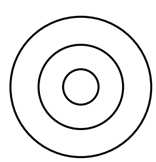
Electron Configurations, Rutherford-Bohr and Electron Dot Diagrams

You should now know three different ways to represent atoms: electron configurations (Grade 11), Rutherford-Bohr diagrams (Grade 9) and electron dot diagrams (Grade 10). While electron configurations are useful to explain why the electrons in an atom are arranged the way they are, they are tedious to write. And, since it is only the outer electrons which participate in chemical reactions, we can ignore all of the electrons in the full inner shells. For simplicity, we will use electron dot diagrams in our studies of bonding.

Remember, when drawing electron dot diagrams:

- Write the symbol for the element and show only the outer (valence) electrons
- The number of valence electrons is equal to the Group Number (in Roman Numerals) for each element
- Follow the convention of only doubling up the electrons after all four "orbitals" have one electron each

eg. the electron dot diagram for phosphorus (Group V) would be $\overset{\bullet}{\underset{\bullet}{:P\cdot}}$

Element	Atomic #	Electron Configuration	Rutherford-Bohr Diagram	# of Valence Electrons	Electron Dot Diagram
Na					
Mg					
O					
Al					
C					
N					

Homework:
After you have completed the above chart, draw the electron dot diagrams for atoms with atomic number 1, 9, 10, 14, 15, 16, 17, 18, 20, 34, 35, 36, 37, 38, 52, 53, 54, 55, 56 and 85

Covalent Bonding

- covalent bonds form between ____________________ and ____________________
- atoms __________ valence electrons to complete a ______________________ electron arrangement

Unlike ionic compounds, covalent compounds form discrete, single particles called ________________. Each molecule contains the number and type of atoms given in its molecular formula, bonded together.

Bonding Capacity: the number of covalent bonds an atom will form (to complete a stable octet).

Elements	Group Number	Number of Valence Electrons it HAS	Number of Valence Electrons it NEEDS	Bonding Capacity
H				
F, Cl, Br, I				
O, S, Se				
N, P				
C, Si				

To draw covalent compounds:

1. Draw the electron dot diagrams for all of the atoms in the molecule. Draw the atom with the highest bonding capacity in the ___________ (this is usually the ________ atom in the chemical formula).
2. Show the electrons being shared between atoms. Only _____________ electrons in each atom are available for bonding. Check that all atoms (except H) are surrounded by a stable octet of electrons.
3. Draw the shared (bonded) electrons with a dash "–". This is called a ________________________. Draw in all unshared electron pairs.

eg. NH_3

eg. CF_2Cl_2

Diatomic Elements

Several elements are found in nature as diatomic molecules, formed when two atoms of the ____________ ______________ are covalently bonded together. ("di" means two)

- There are seven diatomic elements:
- These elements can be remembered with the mnemonic " __________________ "

eg. Cl_2 O_2

Multiple bonds

Elements that have a bonding capacity of ______________________ can form multiple bonds.

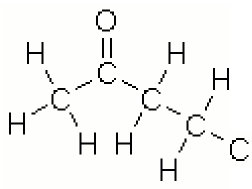
- Double bonds (2 shared pairs of electrons) are stronger than single bonds.
- Triple bonds (3 shared pairs of electrons) are stronger than both single and double bonds.

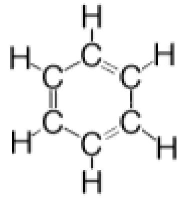
eg. N_2 NOCl

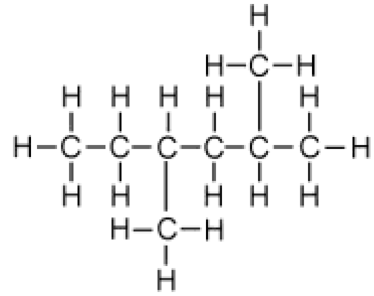
Carbon

Carbon is an unusual element in that it can form bonds with other carbon atoms in __________________ and ________. This ability to bond itself is the basis of _________________________. Carbon needs ___ electrons (it has a bonding capacity of ___). It can form ___, ___, or ___ bonds with other carbon atoms. This means that carbon-based (organic) molecules can be very diverse.

eg.







Two "hints" when you are drawing carbon-based molecules:

- if a molecule contains two or more carbon atoms, they are usually bonded to _____________________ (there may be either ____________ , ____________ or ___________ bonds between the carbon atoms)
- if a molecule contains two (or more) oxygen atoms, they usually are _____ bonded to each other, except for , and .

PCl_3	CH_2O
H_2S_2	CSe_2
C_3H_6	

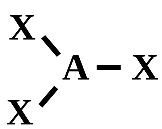
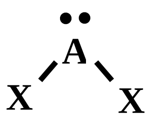
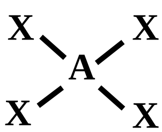
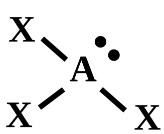
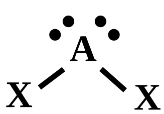
Homework

1. Define octet rule, covalent bond, bonding capacity, molecular formula.
2. Explain how a ***formula unit*** of an ionic compound is different from the ***molecular formula*** of a covalent compound.
3. Draw the Lewis structures (structural formulas) for the following molecules. Be sure to draw in all unshared electron pairs.

   Br_2, CℓF, PH_3, CI_4, P_2H_4, CHCℓ$_3$, H_2S, H_2O_2, C_2H_6, C_2H_4, C_2H_2, CO_2, N_2H_2, N_2, SiO_2, HNO_2, C_3H_8, C_4H_{10}, C_4H_8, C_2H_5OH, HCN, C_6H_{14}, C_3H_7OH, HCOOH, CH_3COOH
   (and a challenge for fun, urea in urine: CON_2H_4 and oxalic acid: $C_2O_4H_2$ found in kidney stones)

Homework: The Shapes of Molecules

1. Complete the chart below using AXE notation (AXnEm) to show the number of bonded electron groups on the central atom (n), number of lone electron pairs (LP) on the central atom (m), the total number of electron groups on the central atom (n+m) and the name of the shape of the molecule.

Drawing of Molecule	**AXE Notation (AXnEm)**	**# of bonded electron groups on the central atom (n)**	**# of lone pairs on the central atom (m)**	**total # of electron groups on central atom (n + m)**	**Name of the Shape of the Molecule**
X — A — X					
					
					
					
					
					

2. For the following molecules:
a) draw the molecule following the octet rule
b) determine the AXE notation for the shape of the molecule
c) name the shape of the molecule

HCN	SiS_2	$COC\ell_2$	NBrO	PH_3
NSF	OF_2	$CC\ell_3F$	AsI_3	H_2Se

Electronegativity and Bonding

Chemical bonds form when atoms __________, ___________ or ____________ their valence electrons to complete a __________________________ (   ) electron arrangement. The type of bonding that will take place between two atoms is determined by the ________________________________ of the atoms in the bond.

Electronegativity is defined as a measure of the strength of attraction that an atom has for the _____________ ____________________. The electronegativity (EN) values for the elements are reported on the ____________ of your Periodic Table in the ___________, ______________ corner.

$_{11}$Na ___________ $_{20}$Ca ___________ $_{16}$S ___________ $_{8}$O ___________ $_{10}$Ne ___________

It is the ______________________ in electronegativity (______) between two bonded atoms that determines whether a bond is ______________ or _____________________, and whether a covalent bond is ________________________ covalent or ________________ covalent.

If the ______________________ between the electronegativities (ΔEN) is:

Less than 0.50:

- The bond is _________________________________
- The electrons are essentially _______________________________ by both atoms
- The bond is ________________________

eg.

Between 0.50 and 1.70:

- The bond is __________________________________
- The electrons are ____ by the atoms
- The electrons are pulled towards the atom with the ____. This means that the atom with the higher electronegativity will have a ________________________________ __________________. The atom with the lower electronegativity will have a __________________ ___________________________________.

eg.

Greater than 1.70:

- The bond is ________________
- The electrons are ____ from the atom with the lower electronegativity (usually a _______________) and are completely transferred to the atom with the ____ (usually a ______________________)
- The metal will have a ____________________________________ (it will be a ________________) and the non-metal will have a ____________________________________ (it will be a ______________________)

eg.

Classifying Bonds and the Bonding Continuum

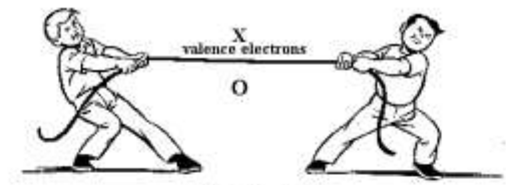
Atoms are most stable when they have a stable octet electron arrangement in their valence shell. One way that atoms fill their valence level is by sharing valence electrons with one or more other atoms. When two atoms are simultaneously attracted to the same electrons, it pulls the atoms together. A pair of valence electrons that is shared between two atoms forms a chemical bond called a **covalent bond**. The sharing of electrons is possible because the positively charged nucleus of each atom is attracting, not only its own electrons, but also the electrons of the other atom.

Electronegativity (EN) is a measure of the strength of the attraction that an atom has for the electrons in a bond. It is reported on the back of the Periodic Table, in the top right hand corner for each element. The larger the value of EN, the more strongly an atom attracts electrons.

The difference in electronegativity (ΔEN) between the bonding atoms affects the nature of a chemical bond.

If the difference in electronegativity (ΔEN) is less than 0.50:

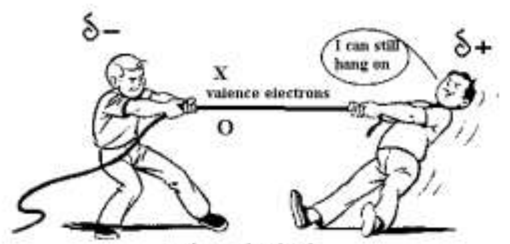
- both atoms have a very similar attraction for the bonding electrons
- the bonding electrons are shared approximately equally between the atoms
- the bond is uncharged
- the bond is a non-polar covalent bond



non-polar covalent bond

If the difference in electronegativity (ΔEN) is 0.50 – 1.70:

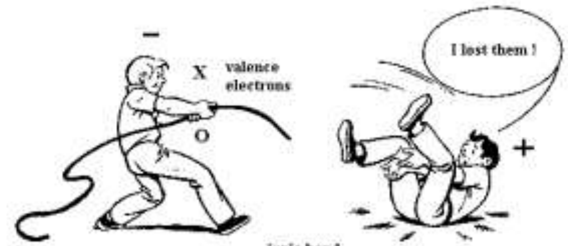
- one of the bonding atoms has a significantly stronger attraction for the electrons than the other atom
- the electrons are shared unequally between the atoms
- the electrons are shifted closer to the atom with higher EN forming a slight negative charge (δ–), and shifted away from the atom with lower EN forming a slight positive charge (δ+)
- the bond is a polar covalent bond



polar covalent bond

If the difference in electronegativity (ΔEN) is more than 1.70:

- one of the bonding atoms has a much, much stronger attraction for the bonding electrons than the other atom
- the electrons are pulled completely away from the atom with lower EN forming a positive ion (+) and transferred to the atom with higher EN, forming a negative ion (–)
- the negative and positive ions are attracted to one another by electrostatic attraction, forming an ionic bond



ionic bond

The difference in electronegativity between bonded atoms can range anywhere from ΔEN = 0.00 (for two identical atoms), up to ΔEN = 3.30 (between atoms of cesium and fluorine). The range of values is called a "continuum". When ΔEN = 0.00, the bond is entirely non-polar (or pure) covalent. As ΔEN increases, the bond becomes increasingly polar, with Cs – F being the most extremely polar bond.

The Bonding Continuum:

EN = 0 — 0.50 — 1.70 — EN = 3.3

ΔEN = 0.00 – 0.50	ΔEN = 0.50 – 1.70	ΔEN = 1.70 – 3.30
non-polar covalent bond	polar covalent bond	ionic bond

increasing bond polarity →

Summary:

ΔEN	Type of Bond	Where the Bonding Electrons are Found	Charge on Bonded Atoms
< 0.50	non-polar covalent	equally shared between bonding atoms	uncharged (no charged regions)
0.50 – 1.70	polar covalent	shifted closer to the atom with higher electronegativity	partial positive (δ+) and partial negative (δ–) charges
> 1.70	ionic	transferred to the atom with higher electronegativity	full positive (+) and full negative (–) charges

Homework:

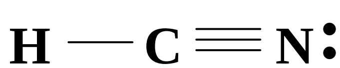
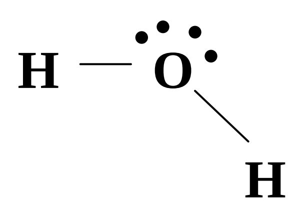
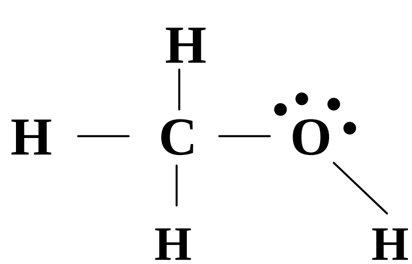
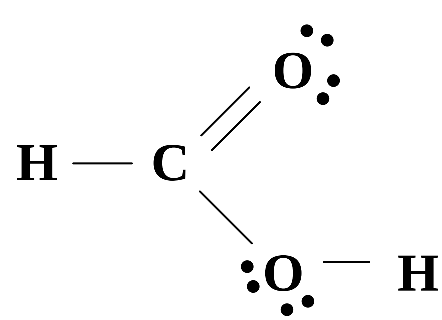
1. Read over the “tug-of-war” note very carefully.
2. For the following bonds:

a) calculate the difference in electronegativity (ΔEN) between the bonding atoms

b) determine the type of bonding that will occur: non-polar covalent, polar covalent, or ionic
   - if the bond is non-polar, it is uncharged so do not draw in any charged regions
   - if the bond is polar, label the appropriate atoms with partial positive (δ+) and partial negative (δ–) charges
   - if the bond is ionic, label the appropriate atoms with full positive (+) and full negative (–) charges

H – O	Na – Br	Mg – O	C – F
H – S	N – O	P – H	B – Cl
K – Cl	C – H	F – Se	S – C

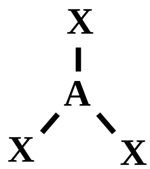
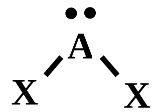
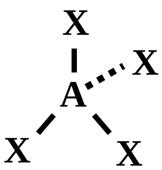
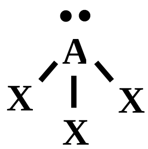
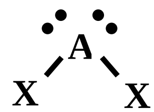
3. Give three examples of bonds for which ΔEN = 0.0.
4. For the following molecules, determine the type of bond(s) and label any partial or full charges.

Hydrogen cyanide (a poisonous gas) HCN	**Dihydrogen Monoxide (water) H_2O**
	
Methanol (a poisonous alcohol) CH_3OH	**Formic Acid (in red ant stings) HCOOH**
	

Determining the Polarity of Molecules

A molecule is ____________ if any of the bonding electrons are ________________ toward one atom in the molecule. To determine the polarity of a molecule, follow these steps:

Step 1: Draw the stick diagram of the covalent molecule using the Octet Rule.
Step 2: Write the AXE notation (AXnEm) for the molecule and name its shape.
Step 3: Evaluate the symmetry of the molecule's shape.

 Shape is: ____________________	 Shape is: ____________________	 Shape is: ____________________
 Shape is: ____________________	 Shape is: ____________________	 Shape is: ____________________

Molecules with the AXE notation ________________ (__________________________ on the central atom) are ______________________ in shape.

Step 4: Evaluate the symmetry of the molecule's bonded atoms
eg. H – C ≡ N: The bonded atoms are: ____________________________
eg. : O = C = O: The bonded atoms are: ____________________________

Step 5: Evaluate the overall symmetry of the molecule considering BOTH its shape and bonding atoms

Step 6: Calculate the ΔEN for all types of bonds:

ΔEN < 0.50 the bond is __________________________
ΔEN 0.50 to 1.70 the bond is __________________________
ΔEN > 1.70 the bond is __________________________

Step 7: Use the overall symmetry and ΔEN to determine the polarity of the overall molecule

- If the molecule is symmetrical (both its shape and bonded atoms) then the molecule is _______________, regardless of the polarity of its bonds because any charges will cancel out. Do not label any charges.
- If the molecule is asymmetrical (either its ____________ or bonded ___________) AND it has non-polar bonds (ΔEN ____________) then the molecule is _____________________________________. The charges are there, but they are very small so do not label _______ and _______ on the molecule.
- If the molecule is asymmetrical (either its ____________ or bonded ___________) AND it has polar covalent bonds (ΔEN __________________) then the molecule is ___________________. Label _______ and _______ on the molecule.

Practice Determining the Polarity of Molecules

CO_2

AXE notation:

Name of Shape:

Symmetry of Shape:

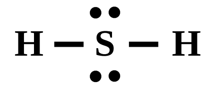
Symmetry of Atoms:

$\triangle$EN

Polarity of Molecule:

Charges? If yes, label them on the molecule.

H_2S



AXE notation:

Name of Shape:

Symmetry of Shape:

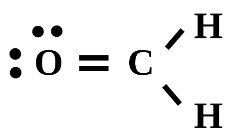
Symmetry of Atoms:

$\triangle$EN

Polarity of Molecule:

Charges? If yes, label them on the molecule.

CH_2O



AXE notation:

Name of Shape:

Symmetry of Shape:

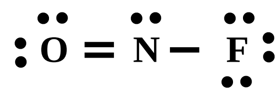
Symmetry of Atoms:

$\triangle$EN

Polarity of Molecule:

Charges? If yes, label them on the molecule.

NOF



AXE notation:

Name of Shape:

Symmetry of Shape:

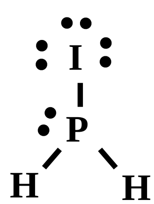
Symmetry of Atoms:

$\triangle$EN

Polarity of Molecule:

Charges? If yes, label them on the molecule.

PH_2I



AXE notation:

Shape:

Symmetry of Shape:

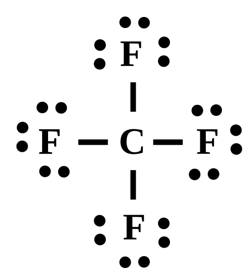
Symmetry of Atoms:

$\triangle$EN

Polarity of Molecule:

Charges? If yes, label them on the molecule.

CF_4



AXE notation:

Shape:

Symmetry of Shape:

Symmetry of Atoms:

$\triangle$EN

Polarity of Molecule:

Charges? If yes, label them on the molecule.

Homework: Draw stick diagrams for the compounds below. Determine the overall polarity of each molecule. Label any charges. NH_3 SF_2 CBr_2I_2 HCN CSF_2 $PC\ell O$ CS_2 NSI $AsC\ell_3$