Unit 3: Periodic Table

Mrs. Snyder
Honors Chem

Unit 1 Learning Objectives:

By the end of the unit students will be able to…

- Locate and state the important properties of the main chemical families including the alkali metals, alkaline earth metals, halogens, noble gases, lanthanides, actinides and transition metals.
- Explain and define the following periodic trends, and how they relate to atomic structure.
    - Atomic Radius
    - Ionization Energy
    - Electronegativity
    - Ionic radius
- Draw Lewis Structures from Chemical Formulas
- Assign bond order
- Calculate the total number of valence electrons in a polyatomic ion
- Draw a Lewis Structure for Polyatomic Ions
- Assign formal charges to atoms and polyatomic ions
- Draw resonance structures for polyatomic ions
- Classify bonds as either ionic or covalent and as either polar or non-polar using electronegativity values.
- Assign shapes to molecules using the VSEPR Theory and draw the VSEPR diagrams for a molecule
- Classify molecules as polar or non-polar using shape
- Compare miscible, and immiscible, by definition and with example and determine based on polarity.

Monday	Tuesday	Wednesday	Thursday	Friday
October 1 Intro: Paint Chip Activity Notes: Periodic Trends	**2** Lab: Periodic Trends	**3** Notes: Lewis Structures	**4** Notes: Lewis Structures	**5** Chemical Families Activity
8 Chemical Families Presentations	**9** Alien Periodic Table Challenge Activity	**10** Notes: Bonding, Polarity and VSEPR Theory	**11** Notes: Bonding, Polarity and VSEPR Theory	**12** Lab: Lewis Structures
15 Review	**16** **UNIT 3 TEST** **Homework Packet Due**	**17**	**18**	**19**

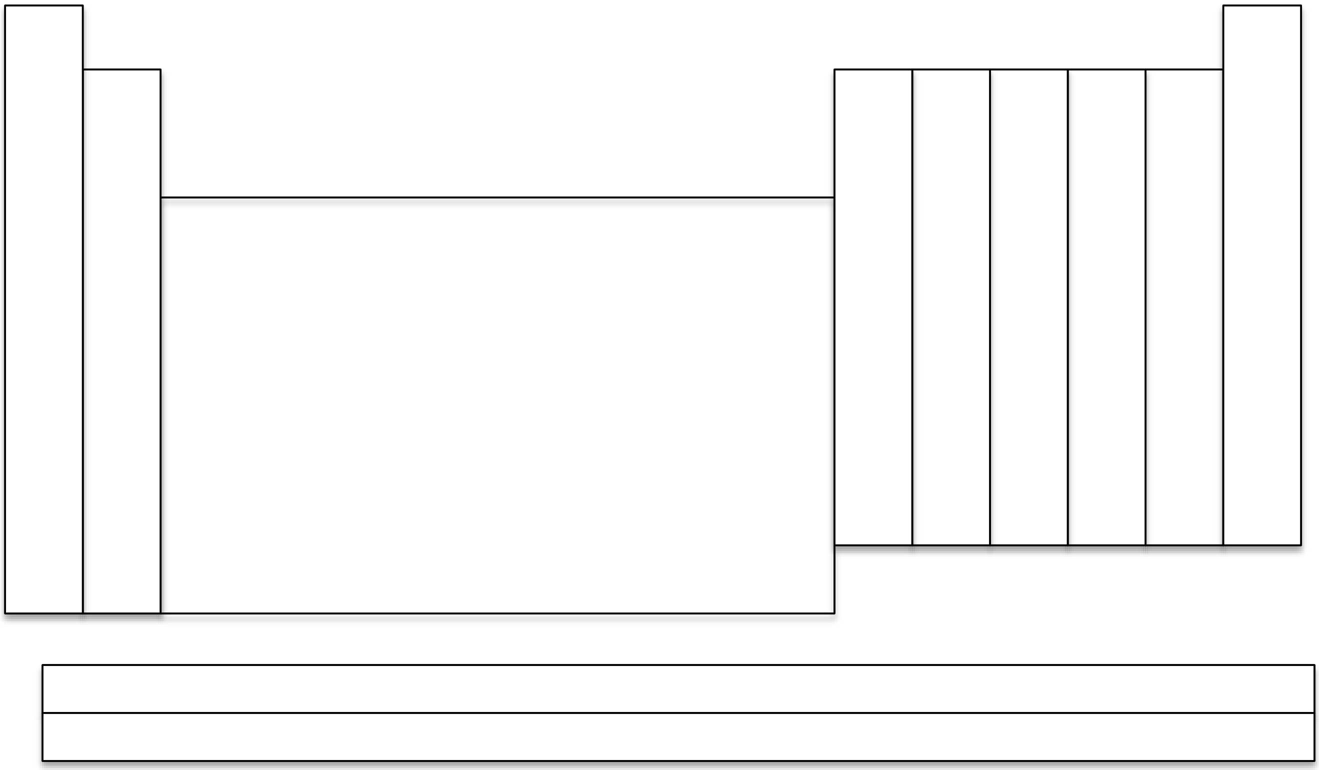
Periodic Table Trends

Periodic Trends: ____

____

Periodic Trend	Definition	What happens when you do down a group?	What happens when you go across a row?	Examples
Atomic Radius				
Ionization Energy				
Electronegativity				
Ionic Radius				

Summary of Periodic Trends



Lewis Structures and Bonding

Electron Dot Structures

Elements can be represented by **electron dot structures** to show the number of valence electrons. Recall, the number of valence electrons is equal to the group number. Only the valence electrons are shown, as these are the electrons that participate in bonding.

Example: Draw the electron dot structure for chlorine

Ex: Draw the electron dot structure for carbon

Atoms will bond with each other in order to ____. In general, atoms will follow the ________________ rule and share ___________________ so that they have ______ electrons in their valence shell. *Note: There are exceptions to this rule.

Lewis Structure: ____

Example: Draw a Lewis Structure for the following compounds. The subscript in the formula indicates the number of atoms of each element in the compound. (ie, in Cl_2 the subscript "2" indicates that there are two chlorine atoms"

Rules for Drawing Lewis Structure:

Step 1: Add up the number of valence electrons in all of the atoms.

Step 2: Write down the most likely arrangements.

Central atom usually comes 1st in chemical equation (exception acids)

The central atom is the atom with the lowest ionization

Arrange the atoms symmetrically around the central atom

Step 3: Place one electron pair between each pair of bonded atoms.

Step 4: Complete the octets or duplet (H). If there are not enough electron pairs, form multiple bonds.

Step 5: Represent each bonded electron pair by a line.

Cl_2	SBr_2
CI_4	SiO_2
AsF_3	HCN Note: Hydrogen has a full valence shell with only two electrons. H C N

Lone Pairs: ____

Bond Order: ____

Bond Order =1 for a single bond, two electrons are shared
Bond Order= 2 for a double bond, four electrons are shared
Bond Order=3 for a triple bond, six electrons are shared
Bond Order = 4 for a quadruple bond, eight electrons are shared.
* A compound my have more than one bond order. In this case, each type of bond is labeled separately.

Exceptions to the Octet Rule (Exceptions to having a full valence shell)

Some elements can be stable in a compound when they do not have a full valence shell of eight electrons.

Electron Deficient: valence shell is unfilled with fewer than eight electrons.
Ex. Boron can be stable with only six electrons in the valence shell

Ex. BF_3

Expanded Octet: valence shell is over filled with more than eight electrons.
Ex. Phosphorus can be stable with ten electrons in the valence shell.

Ex. PCl_5

Ex. Sulfur can be stable with twelve electrons in the valence shell

Ex. SF_6

Polyatomic IONS

Polyatomic ions are ____

When drawing the Lewis Structure for a polyatomic ion, add one electron for each negative charge and subtract one electron for each positive charge.

The charge on each atom in the compound can be determined by calculating the **formal charge**. The formal charges for all atoms must add up to the overall charge for the polyatomic ion.

Formal Charge
The formal charge on each atom in a polyatomic ion can be calculated as follows:

Formal Charge = Number of valence electrons – number of unshared electrons – $\frac{1}{2}$ shared electrons

Practice: Draw a Lewis Structure for the following polyatomic ions. Determine the total number of valence elections in the ion and calculate the formal charge on each atom.

Ex. Ammonium NH_4^+

Ex. Nitrate NO_3^-

Resonance Structures: ____

____

Ex. NO_3^-

Bonding and Polarity

Bond Type and Bond Polarity depend on the relative electronegativity values of atoms.

Electronegativity ____

Electronegativity (or the difference in electronegativity) helps to classify the bond type and bond polarity. To determine the different in electronegativity subtract the electronegativity values of the atoms in a bond.

Electronegativity Values

H 2.1																	He 0
Li 1.0	Be 1.5											B 2.0	C 2.5	N 3.0	O 3.5	F 4.0	Ne 0
Na 0.9	Mg 1.2											Al 1.5	Si 1.8	P 2.1	S 2.5	Cl 3.0	Ar 0
K 0.8	Ca 1.0	Sc 1.3	Ti 1.5	V 1.6	Cr 1.6	Mn 1.5	Fe 1.8	Co 1.9	Ni 1.8	Cu 1.9	Zn 1.6	Ga 1.6	Ge 1.8	As 2.0	Se 2.4	Br 2.8	Kr 0
Rb 0.8	Sr 1.0	Y 1.2	Zr 1.4	Nb 1.6	Mo 1.8	Tc 1.9	Ru 2.2	Rh 2.2	Pd 2.2	Ag 1.9	Cd 1.7	In 1.7	Sn 1.8	Sb 1.9	Te 2.1	I 2.5	Xe 0
Cs 0.7	Ba 0.9	Lu 1.2	Hf 1.3	Ta 1.5	W 1.7	Re 1.9	Os 2.2	Ir 2.2	Pt 2.2	Au 2.4	Hg 1.9	Tl 1.8	Pb 1.9	Bi 1.9	Po 2.0	At 2.2	Rn 0
Fr 0.7	Ra 0.9																

Covalent Bonds:

If the difference in electronegativity between two atoms in a bond is small (between 0 and 1.7), the bond will be *covalent. Since the atoms are comparable to one another in terms of electronegativity, they each have a similar pull on the electrons and will SHARE the electrons.*

Ionic Bonds:

If the difference in electronegativity between to atoms in a bond is large (1.8 or greater), the bond will be *ionic. Electrons will be TRANSFERRED from the less electronegative atom to the more electronegative atom resulting in the formation of an ion. The attraction of the two ions forms an ionic bond.*

Non-Polar Bonds:

If the difference in electronegativity is between 0 and 0.4, the bond is classified as *non-polar*. *The electrons are equally shared between the atoms in a bond.*

Polar Bonds:

If the difference in electronegativity is 0.5 or greater, the bond is classified as *polar*.

- All ionic bonds are polar since the electrons are transferred and each atom carries a full charge
- For a covalent bond, the electrons are unequally shared and are drawn from the less electronegative element towards the more electronegative element. The more

electronegative atom will carry a partial negative, and the less electronegative will carry a partial positive charge.

Summary of Bonding and Polarity

Bond Type	How are electrons distributed?	What is the difference in electronegativity?
Covalent		
Ionic		

Bond Polarity	How are electrons distributed?	What is the difference in electronegativity?
Non-Polar		
Polar		

Distribution of Charge:

- For a **non-polar covalent** bond there is no distribution of charge
- For an **ionic bond** (all polar), the atom that gains electrons will be labeled with a negative charge, and the atom that loses electrons will be labeled with a positive charge.
- For a **polar covalent** bond an arrow ( ) will be drawn towards the more electronegative atom to show that the electrons are pulled towards that atom. The more electronegative atom will be labeled as partially negative ( ) and the less electronegative atom will be labeled sas partially positive ( )

Practice: Calculate the change in electronegativity (ΔEN), and classify the following bonds as covalent, or ionic and as non-polar or polar. Draw the bond and indicate the distribution of charge.

Ex. F_2

Ex. NaCl

Ex. HBr

VSEPR Theory: Valence Shell Electron Pair Repulsion Theory

To determine the polarity of a molecule, the three-dimensional shape of the molecule must be considered. VSEPR theory predicts the three-dimensional shape of a molecule based on the number of atoms bonded to the central atom and the number of lone pairs on the central atom. The bonds and lone pairs around the central atom will be oriented as far as possible from each other to minimize electron repulsion.

Example Molecule	**Lewis Structure**	**# of atoms bonded to the central atom**	**# lone pairs bonded to central atom**	**Shape**	**VSEPR Diagram**	**Polar or Non-Polar**
CO_2		2	0	Linear		Non-polar
BH_3		3	0	Trigonal Planar		Non-Polar
CH_4		4	0	Tetrahedra l		Non-Polar
NH_3		3	1	Trigonal Pyrimidal		Polar
H_2O		2	2	Bent		Polar
PCl_5		5	0	Trigonal Bipyramid al		Non-Polar
SF_6		6	0	Octahedral		Non-Polar

Symbols Used in VSEPR Diagrams

Polarity of Molecules based on the VSEPR shape

The VSEPR shape can be used to determine polarity of a molecule. A molecule with a "symmetrical" three-dimensional shape will be non-polar. A molecule with an "asymmetrical" three-dimensional shape will be polar.

Substances can be tested for polarity by mixing them with liquids of known polarity. Generally, "like dissolves like" means that ____

____

Miscible: ____

Immiscible: ____

Practice: Draw the Lewis Structure and using VSEPR theory determine the shape of the structure.

Ex. CF_4

Ex. PCl_3

Ex. CS_2