CHEMICAL BONDING

Learning Objectives:

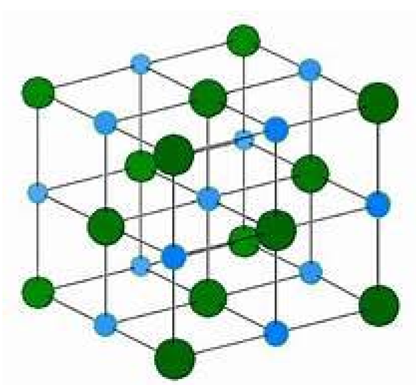
I. Bonding: Basic features of ionic and covalent bonds.
II. Use of Periodic table be used to determine whether an atom forms a cation or an ion, and its resulting ionic charge.
III. Formation of Ionic Compounds
IV. Determination of the formula of an ionic compounds
V. Nomenclature of Ionic Compounds
VI. Properties of ionic compounds.
VII. polyatomic ions
VIII. Useful consumer products and drugs that are composed of ionic compounds.
IX. Formation of Covalent Compounds
X. Nomenclature of Covalent Compounds
XI. Lewis Dot Structures
XII. Resonance
XIII. Electronegativity and bond polarity
XIV. Molecular Shape(VSEPR Theory)
XV. Molecular Polarity
XVI. Intermolecular forces between Covalent molecules

I. Bonding: Basic features of ionic and covalent bonds

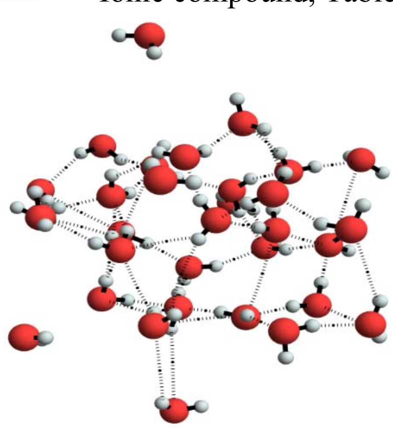
In previous chapters we have seen that compounds are tightly attracted group of atoms. In this chapter we will find how the compounds are formed. Compounds are formed due to strong force of attraction. In chemical world it is described as Chemical Bonding. There are two main types of bonding that we are going to discuss: ionic bonding and covalent bonding. We will examine ionic bonding first.

Both ionic bonding and covalent bonding follow one general rule: Electrons gain, lose or share valence lectrons to attain the lectronic configuration of the noble gas closest to them in the perioidic table which is popularly called octet rule.

Ionic bonds results from the transfer of electrons from one element to another. Ionic bonds form between a metal and a nonmetal .Ionic compound consists of oppositely charged ions that strong electrostatic attraction for each other. Whereas, covalent bonds result from the sharing of electrons between two atoms. Covalent bonds occur between two nonmetals or when a metalloid combines with a nonmetals. Covalent compound form discrete molecules where they are connected by weak intermolecular forces.



Ionic compound, Table Salt (NaCl)



Covalent Compound, water (H_2O)

Questions

Which pairs of elements are likely to form ionic bonds and which pairs are likely to form covalent bonds? a) carbon and hydrogen, b) two bromine atoms, c) sodium and sulfur, d) magnesium and bromine

Ans:

a. covalent bond
b. covalent bond
c. Ionic bond
d. Ionic bond

II. Octet Rule

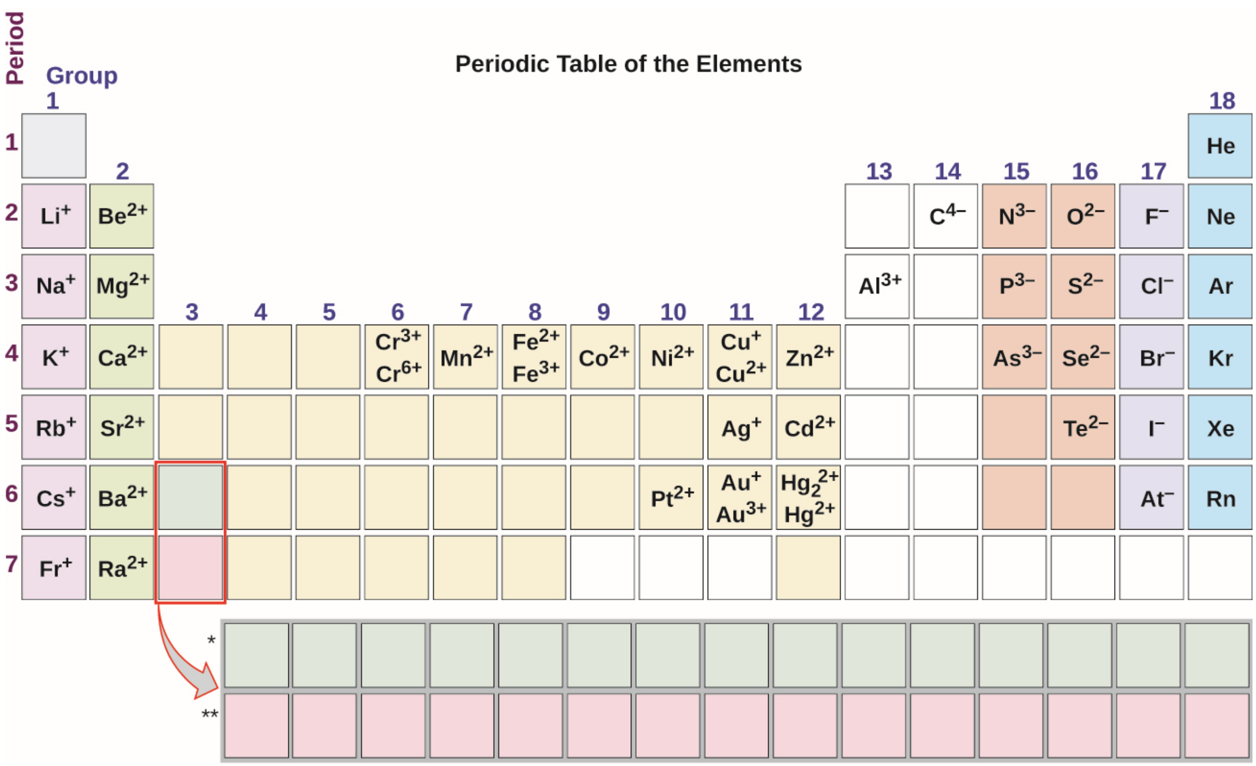
In 1916, Gilbert Lewis proposed a model that explained many observed facts about chemical bonding and chemical reactions. He pointed out that noble gases extremely stable and unreactive due to their eight electronic configuration.

Element	Atomic Number	Electronic Configuration	Group Number	Period Number
Helium	2	$1s^2$	18	1
Neon	10	$1s^22s^22p^6$	18	2
Argon	18	$1s^22s^22p^63s^2\ 3p^6$	18	3
Krypton	36	$1s^22s^22p^63s^23p^63d^{10}4s^2\ 4p^6$	18	4
Xenon	54	$1s^22s^22p^63s^23p^63d^{10}4s^2\ 4p^64d^{10}5s^2\ 5p^6$	18	5
Radon	86	$1s^22s^22p^63s^23p^63d^{10}4s^2\ 4p^6\ 4d^{10}4f^{14}5s^2\ 5p^65d^{10}6s^2\ 6p^6$	18	6

The **octet rule** refers to the tendency of atoms to prefer to have eight electrons in the valence shell. When atoms have fewer than eight electrons, they tend to react and form more stable compounds. When discussing the **octet rule**, we do not consider d or f electrons.
Periodic chart helps us to identify the number of valence electrons of all the elements. On the periodic table Group IA, IIA and IIIA are main group elements and contain metals only. Since metals are located on the left hand side of the periodic chart, metals usually have one two or three valence electrons. Metals tend to lose those electrons to gain previous noble gas configurations. Similarly, main group elements VA, VIA and VIIA elements are nonmetals only. Nonmetals gain extra valence electrons and achieves the next noble gas configuration in that period. Therefore, metals lose electrons equal to their valence

electrons and form positively charged species. Nonmetals gain electrons and form negatively charged species. Positively charged substances are called **cations** and negatively charged substances are called **anions**. By gaining or losing one, two or three electrons an atom forms an **ion** with a completely filled outer shell of electrons.

The charge on main group ions can be predicted from the position in the periodic table. For metals in group 1A, IIA and IIIA, the group number=charge on the cation. For nonmetals in group 5A, 6A and 7A, anion charge =8-group number. Main group elements are specially stable when they possess an octet of electrons. The charge of the ion is written as superscript. Find below the complete charge table for all the metals and nonmetals on the periodic table. Main group elements are specially stable when they possess an octet of electrons. Hydrogen and Boron are exception to the octet rule. Hydrogen needs 2 electrons and boron requires minimum six electrons to satisfy.



Ions are written as Lewis symbol by showing their valence electrons as follows.

Questions

1. Which ions are likely to form?

a. S^-

b. S^{2-}

c. S_3^{3-}
d. Na^+
e. Na_2^{2+}
f. Na^-

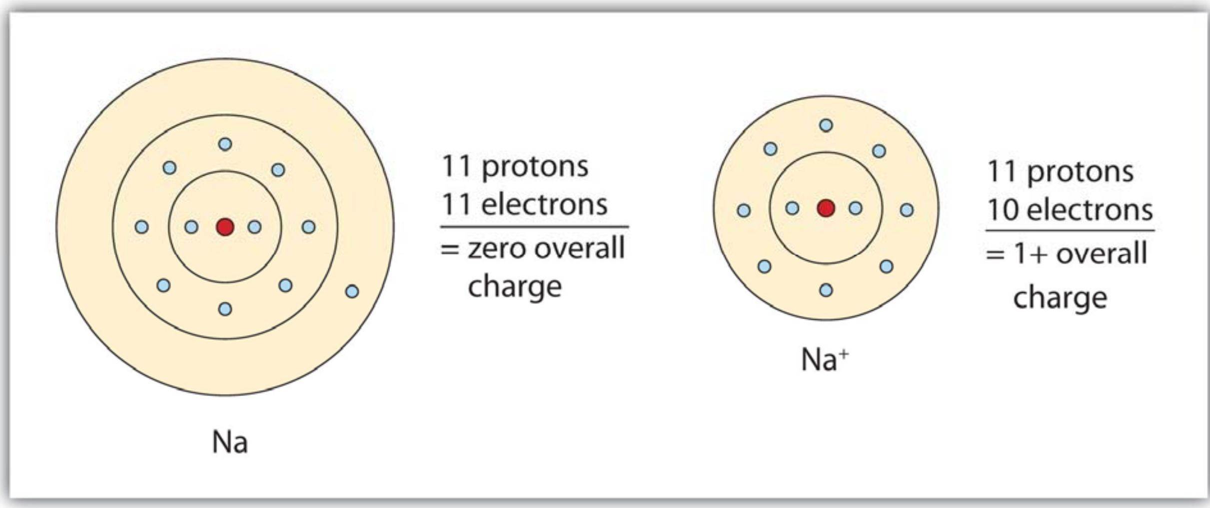
2. What happens when O(oxygen atom) becomes O^{2-}

a) It loses 2 electrons
b) It gains 2 electrons

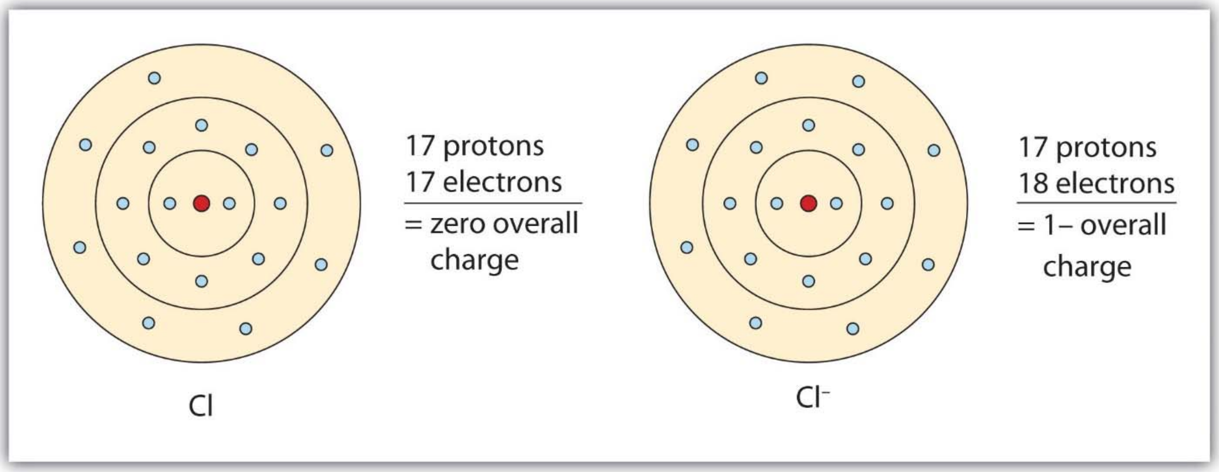
Ans: 1. b, S^{2-} d Na^+ 2. b) it gains 2 electrons

III. Formation of Ionic Compounds

We will discuss the formation NaCl ionic compounds. Ionic compounds are formed between a metal and a nonmetal. Sodium for example is located under Group I. Therefore following octet rule, it will lose one electron to achieve previous noble gas configuration which is equivalent to Neon [2s22p6] and the charge of the sodium ion will be +1.

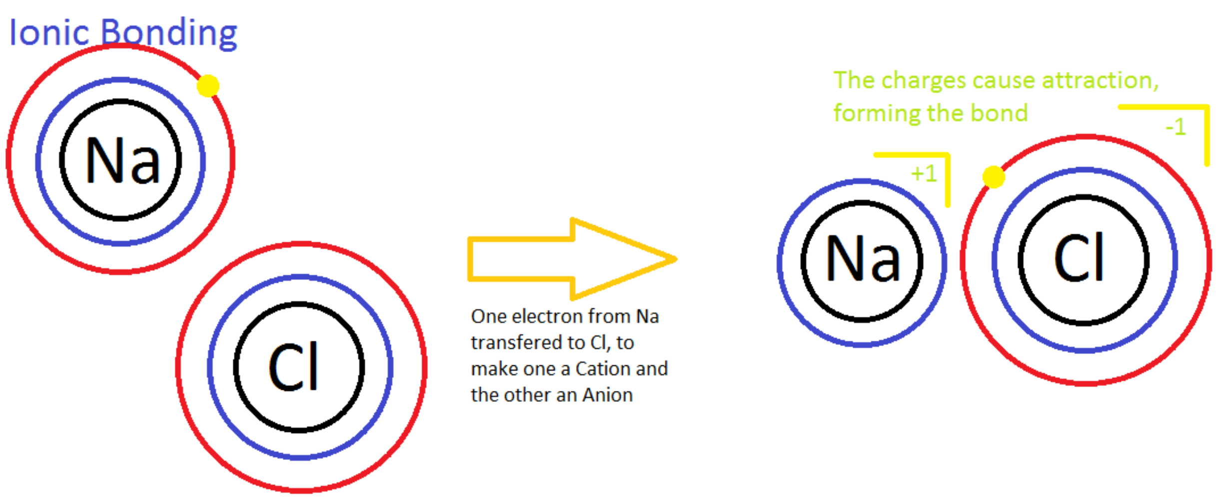


Similarly, chlorine a nonmetal is located under main group seven. Therefore it has seven valence electrons. Chlorine tens to gain one more valence electronsto achieve next noble gas configuration of Ar [3s23p6]. The charge of the chlorine ion is -1.

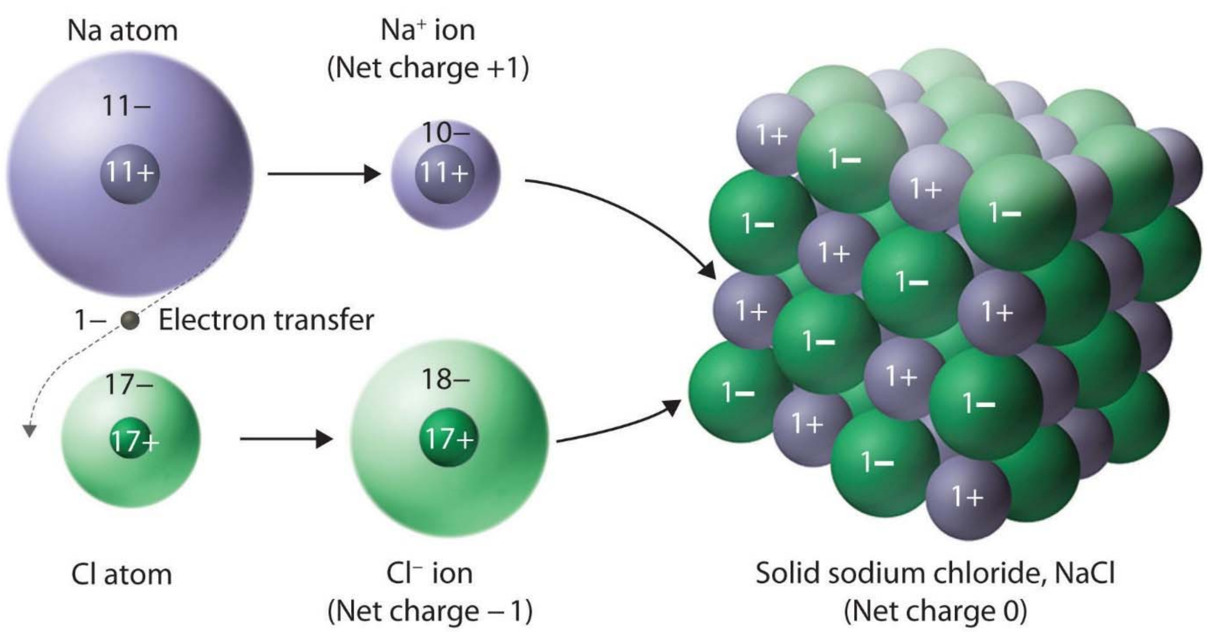


$$e^- + Cl \longrightarrow \underbrace{Cl^-}_{1s^2\,2s^2\,2p^6\,3s^2\,3p^6}$$

When the sodium metal is brought in close contact with chlorine nonmetal, sodium loses electron, chlorine gains electrons and forming ionic bonds due to the electrostatic attraction between the cation and the anion formed Therefore in an ionic compound total charge must be zero.



Following diagram may help to understand the NaCl crystal structure.



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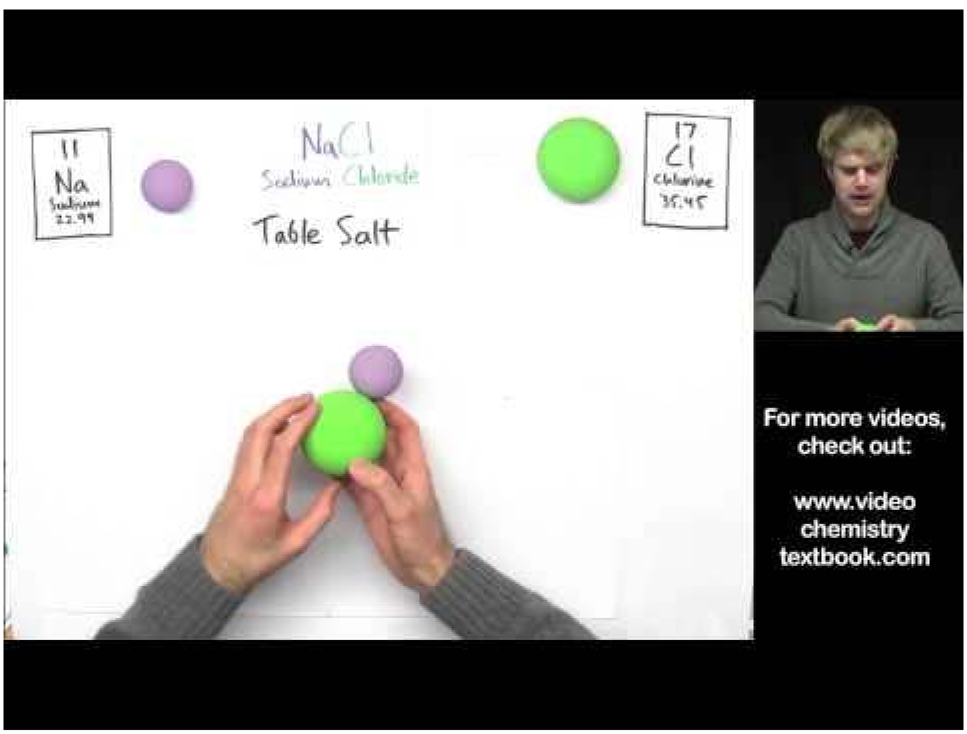
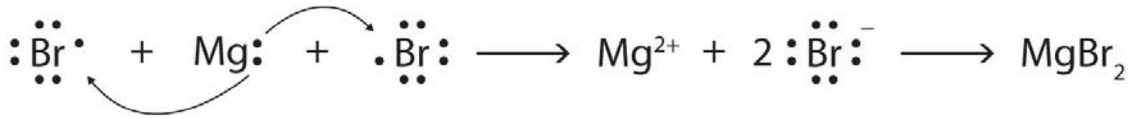
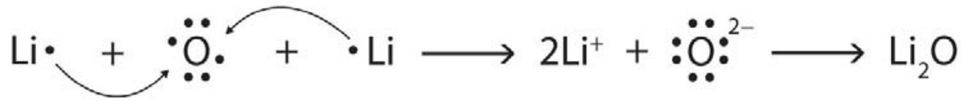


IV. Formula of an ionic compound

Cations and anions always form ionic compounds that have zero overall charge.

Ionic compound s are written with the cation first, and then the anion with subscript to show how many of each are needed to have zero net charge. Ionic compounds are always written with lowest possible ratio between the atoms to form neutral compound.





In short, when the charges are unequal and opposite, ionic compound is formed by switching the opposite charges. When charges are equal and opposite, ionic compounds are formed by simply writing metal followed by nonmetal symbol.

Example: $Mg^{2+} + S^{2-} \rightarrow MgS$
$Al^{3+} + Cl^{-} \rightarrow AlCl_3$
$Pb^{2+} + N^{3-} \rightarrow Pb_3N_2$
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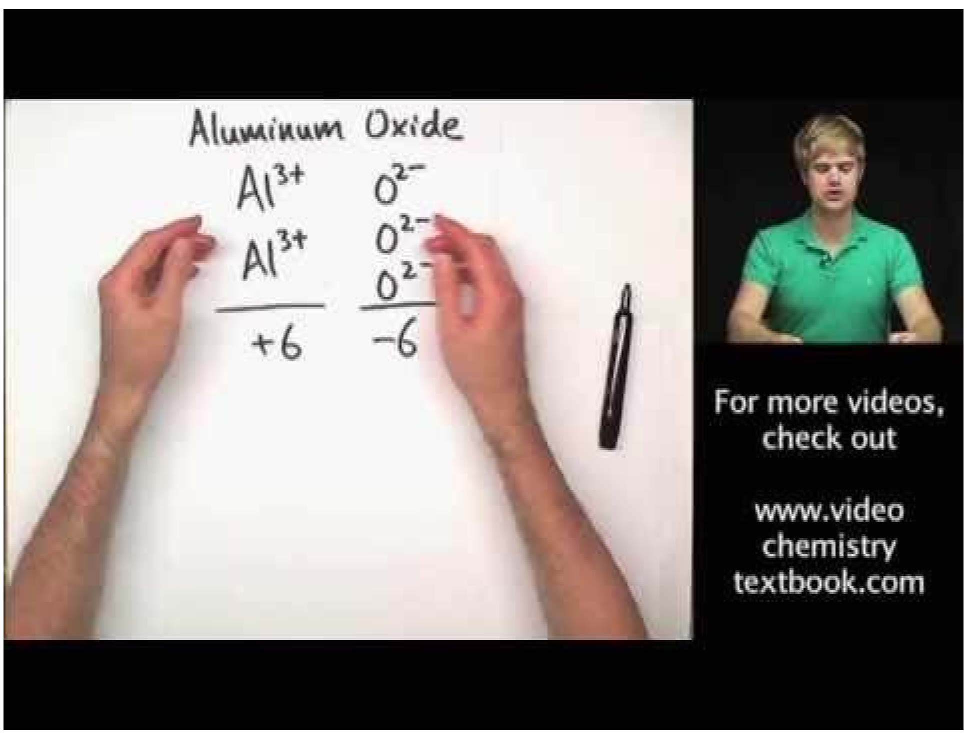


Questions

1. Write the formula for the ionic compound formed from each pair of elements a) Barium and Bromine b) aluminum and sulfur c) zinc and sulfur d) magnesium and fluorine

Ans: a) $BaBr_2$ b) Al_2S_3, c) ZnS, d) MgF_2

V. Nomenclature of Binary Ionic Compounds

Name of monoatomic cations:

When main group metal loses electron(s) based on octet rule, it forms monoatomic cations. The charge of the metal is not required to include in the name because charge is fixed. Here are some examples of common monoatomic ions

Mg^{2+} : Magnesium ion
Al^{3+}: Aluminum ion
Ba^{2+}: Barium ion

Name of the monoatomic anions:

Some names of common monoatomic ions including polyatomic ions with end name -ide. Some common names are given below in the table.

Common Anions

Charge	Formula	Name
1−	H^-	Hydride ion
	F^-	Fluoride ion
	Cl^-	Chloride ion
	Br^-	Bromide ion
	I^-	Iodide ion
	CN^-	Cyanide ion
	OH^-	Hydroxide ion
2−	O^{2-}	Oxide ion
	O_2^{2-}	Peroxide ion
	S^{2-}	Sulfide ion
3−	N^{3-}	Nitride ion

Example: When Ba^{2+} combines with O^{2-}, total charge is neutralized by 1:1 ratio. So the formula of the compound is BaO. And the name: Barium Oxide.
Similarly: $Na^+ + P^{3-} \rightarrow Na_3P$. name: Sodium Phosphide.

In case of transition element, metal has a variable charge, use the overall anion charge to determine the charge on the cation. Then name the cation with charge using a roman numeral with parentheses. For example, in $CuCl_2$, charge of each chloride ion is -1. Therefore, total negative charge is -1*2 for two ions in the formula. In order to make total charge zero, the charge of Cu must be +2. The name of the compound Copper (II) Chloride.
Here are some other examples given below:

formula	systematic name	common name
$CuCl$	copper(I) chloride	cuprous chloride
$CuCl_2$	copper(II) chloride	cupric chloride
Hg_2Cl_2	mercury(I) chloride	mercurous chloride
HgO	mercury(II) oxide	mercuric oxide
FeS	iron(II) sulfide	ferrous sulfide
Fe_2S_3	iron(III) sulfide	ferric sulfide

Common names have been traditionally used for transition metals but systematic names are accepted and used worldwide.

Ionic compounds are always named with the name of the cation first with cations having a fixed charge, the cations have the same name as the name of the original element. The name of the anion usually ends with suffic -ide. It is derived from a single atom or -ate if it is polyatomic.

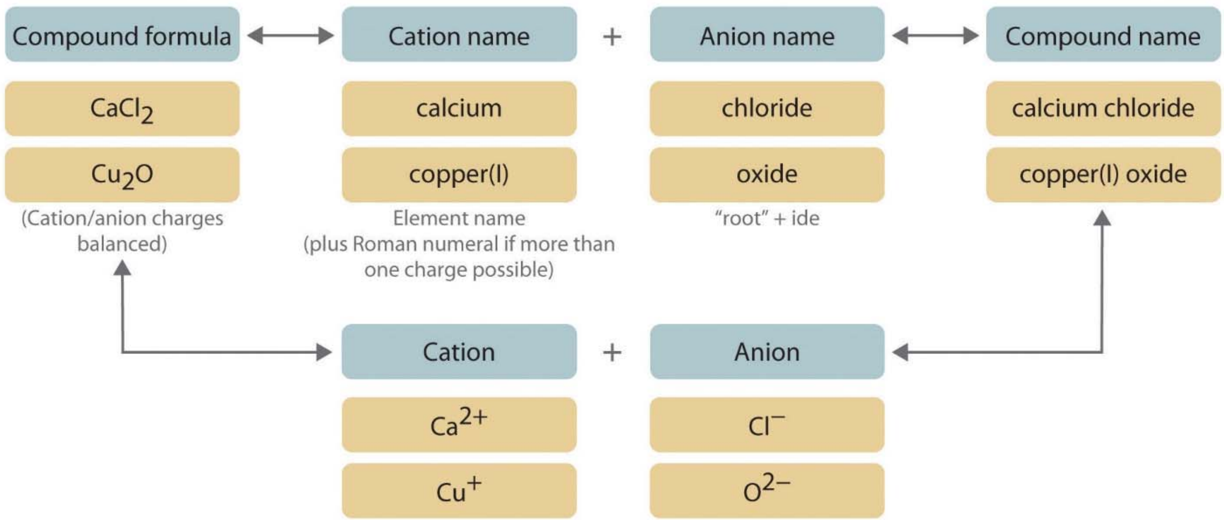
Example: When Ba^{2+} combines with O2- , total charge is neutralized by 1:1 ratio. So the formula of the compound is BaO.

Determination of charge in transition metal ionic compound:
Example: $FeCl_3$

Charge of the metal cation is determined with respect to anion's charge.
In the above example Cl- has charge-1. Since there are total 3 Cl- ions are present in the compound, total negative charge is 3-. In an ionic compound total charge must be balanced. Therefore the charge of the metal cation is 3+
Name of the compound Iron(III) Chloride.

Another example: Co_2O_3
three oxide ions containing 2- each charge make total negative charge -2*3= -6.
There are two metal Co to balance the total charge by +6.
So the charge of one Co metal is +6/2= +3.

The name of the compound Cobalt(III) Oxide.



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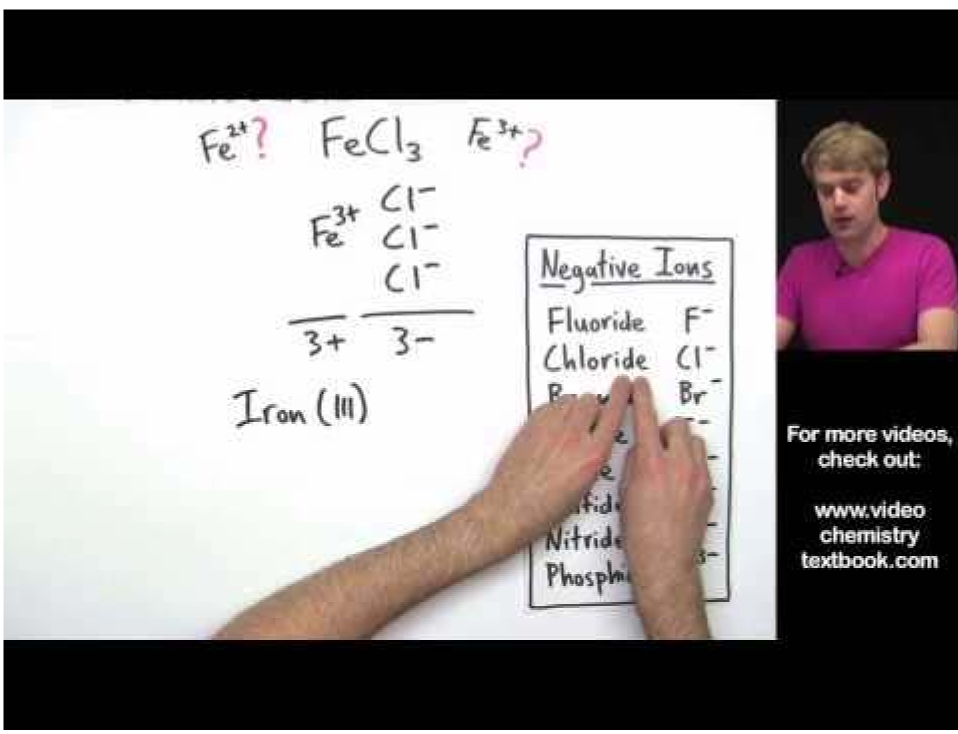
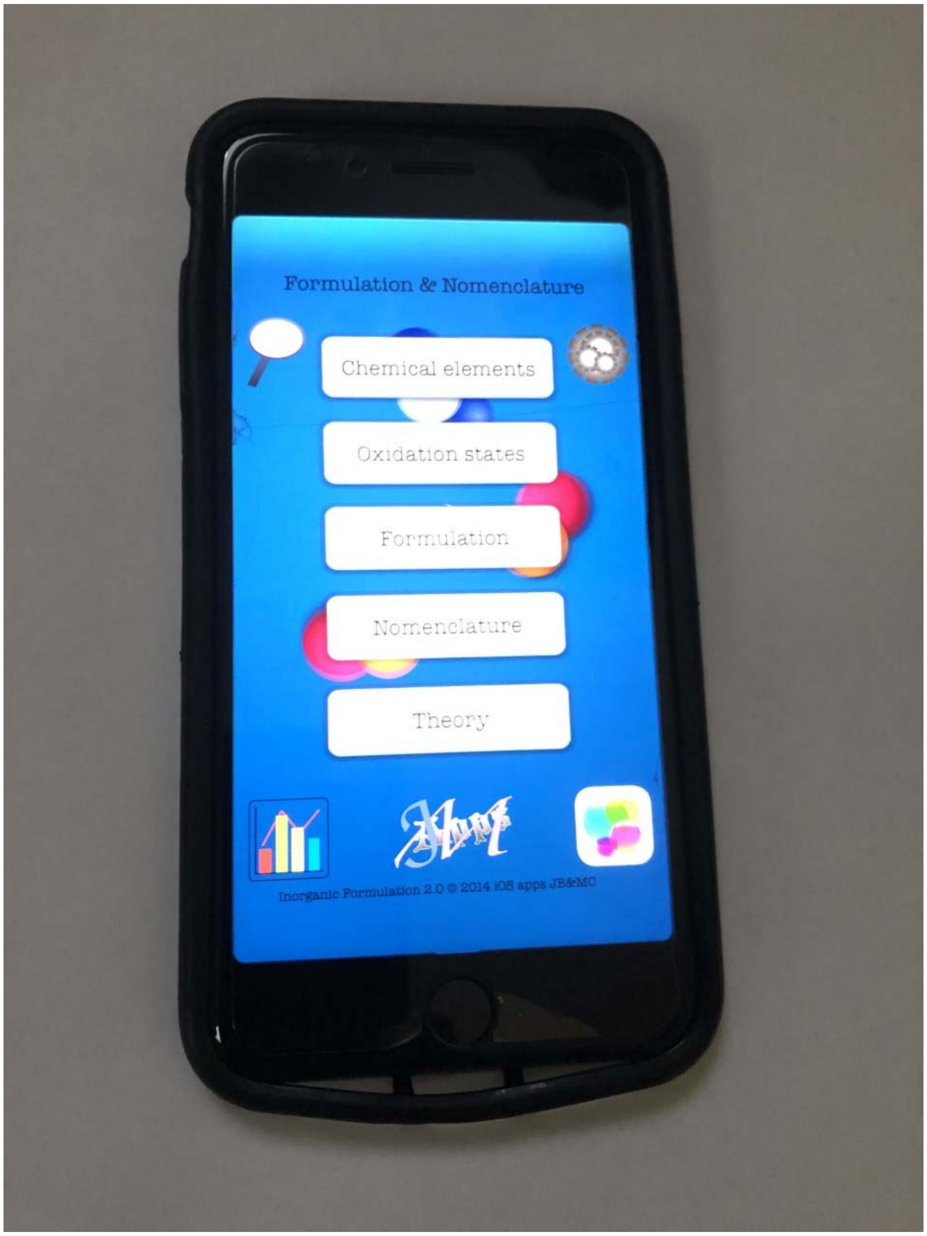


An

<u>An app is available for mobile devices with the name chemicals F& N.</u>
<u>(Formation and Nomenclature) You many practice nomenclature of ionic compound using this app.</u>



Questions

1. Name each ionic compound a. $BaCl_2$, b) $CoBr_3$ c) Cus d) Al_2O_3

Ans: a. Barium Chloride b. Cobalt(III) Bromide c. Copper(II) sulfide, d. Aluminum Oxide

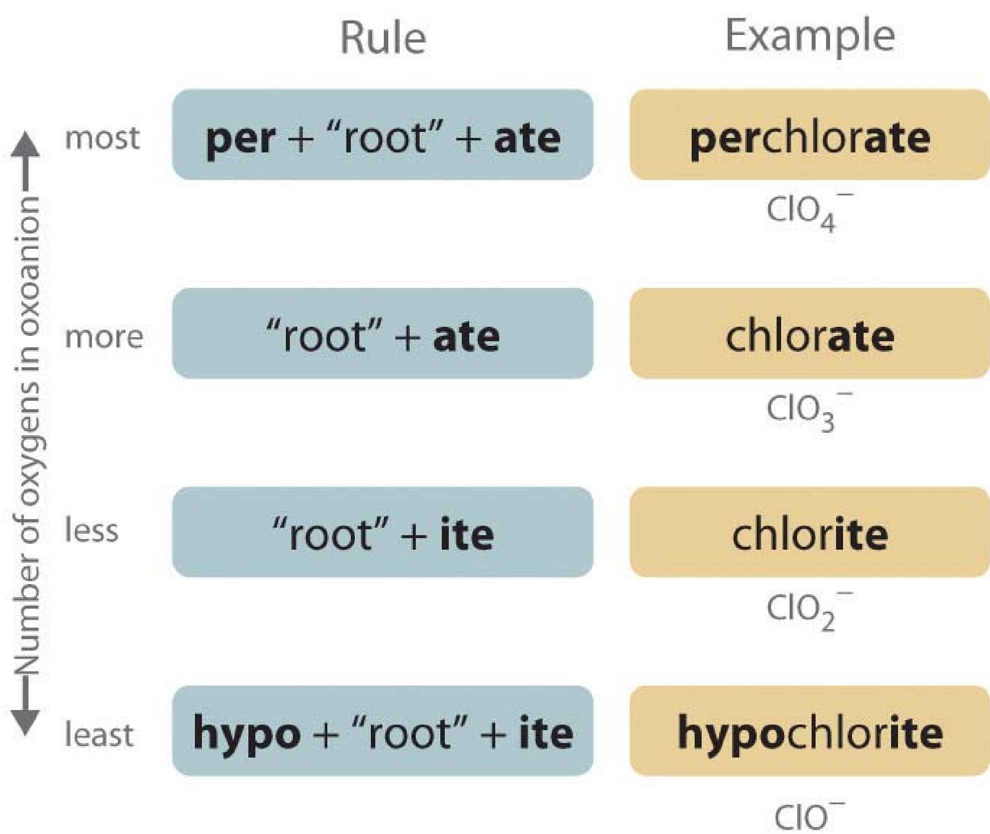
VI. Polyatomic Ions

Polyatomic ions are charged species, consisting of group of atoms. They are composed on more than one element. In polyatomic ion, one central nonmetal attached to either oxygen or hydrogen. Central nonmetals are mainly from period 2 and 3 like N, S, P, Cl. Some metalloids also form polyatomic ions.

The names of polyatomic cations end in the suffix -onium. Many polyatomic anions have names that end in the suffix-ate or ite.
Only one positively charged or polyatomic cation is NH_4^+.
Only two polyatomic anions that have end name -ide (OH^-)are hydroxide and cyanide (CN^-)

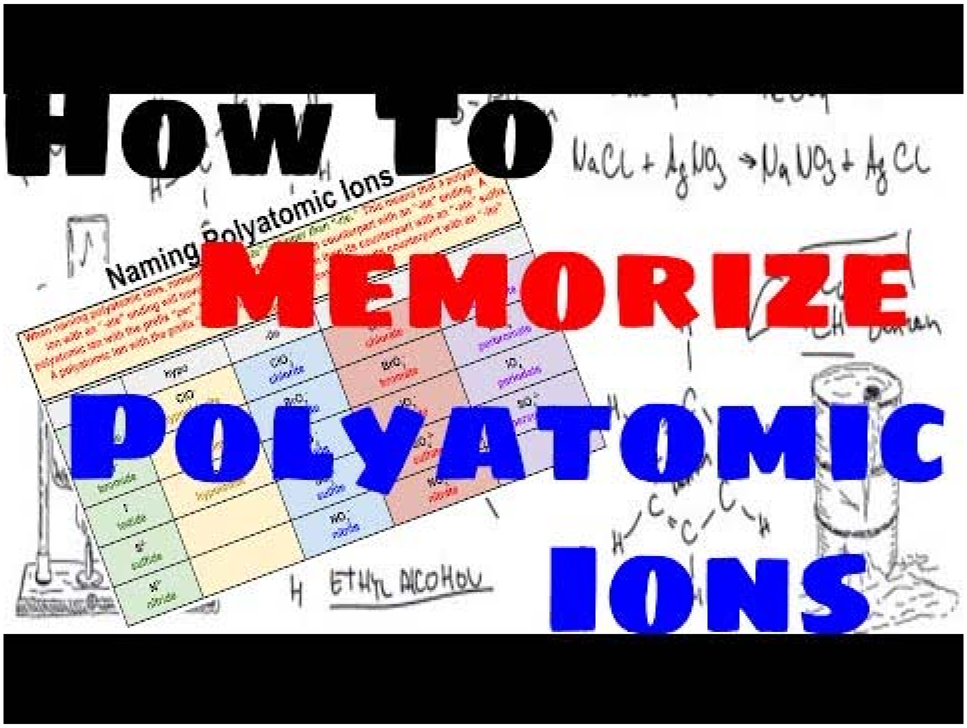
Most polyatomic ions contain oxygen. That is why sometimes they are called oxo anion. Oxo anions have end name either r -ate or -ite. Lower number of oxygen containing polyatomic ion is named as -ite and higher number of oxygen containing polyatomic ions are named as -ate. Example: NO_2^-: nitrite, NO_3^- : nitrate

When a H^+ is added to a polyatomic ion, the negative charge of the ion is decreased by one unit. Example: PO_4^{3-} : phosphate ion, HPO_4^{2-}: hydrogen phosphate ion



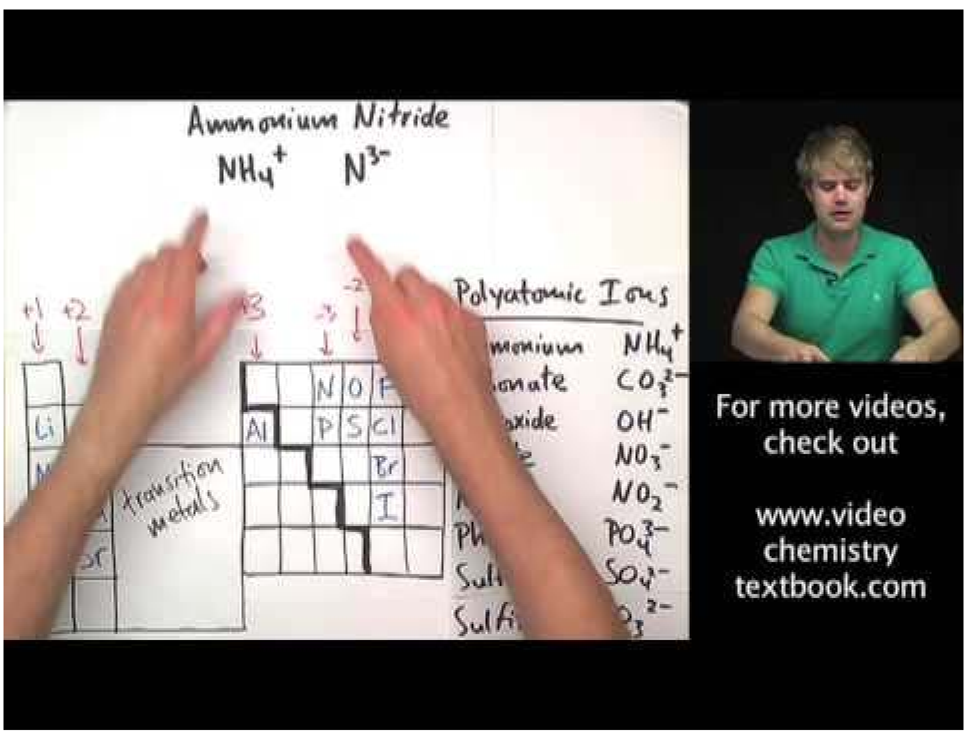
The following video might help to memorize the polyatomic molecules.

https://www.youtube.com/watch?v=zSqfwnkIkvQ&t=86s



The following video explains in detail how to write ionic compounds with polyatomic ions.

https://www.youtube.com/watch?v=p9iQ5Qn42DM&t=5s



Cations

+1
Ammonium, NH_4^+

Anions

-1	-2	-3
Hypochlorite, ClO^- Chlorite, ClO_2^- Chlorate, ClO_3^- Perchlorate, ClO_4^-	Sulfite, SO_3^{2-} Sulfate, SO_4^{2-}	Phosphate, PO_4^{3-}
Nitrite, NO_2^- Nitrate, NO_3^- Bicarbonate, HCO_3^-	Carbonate, CO_3^{2-}	
Hydroxide, OH^-	Peroxide, O_2^{2-}	
Acetate, $C_2H_3O_2^-$	Oxalate, $C_2O_4^{2-}$ Silicate, SiO_3^{2-} Thiosulfate, $S_2O_3^{2-}$	
Permanganate, MnO_4^-	Chromate, CrO_4^{2-} Dichromate, $Cr_2O_7^{2-}$	
Cyanide, CN^- Thiocyanate, SCN^-		

When cations combine with polyatomic ions ionic compounds are formed. When the charges are switched to form neutral compound, parenthesis is used to use for polyatomic ions.

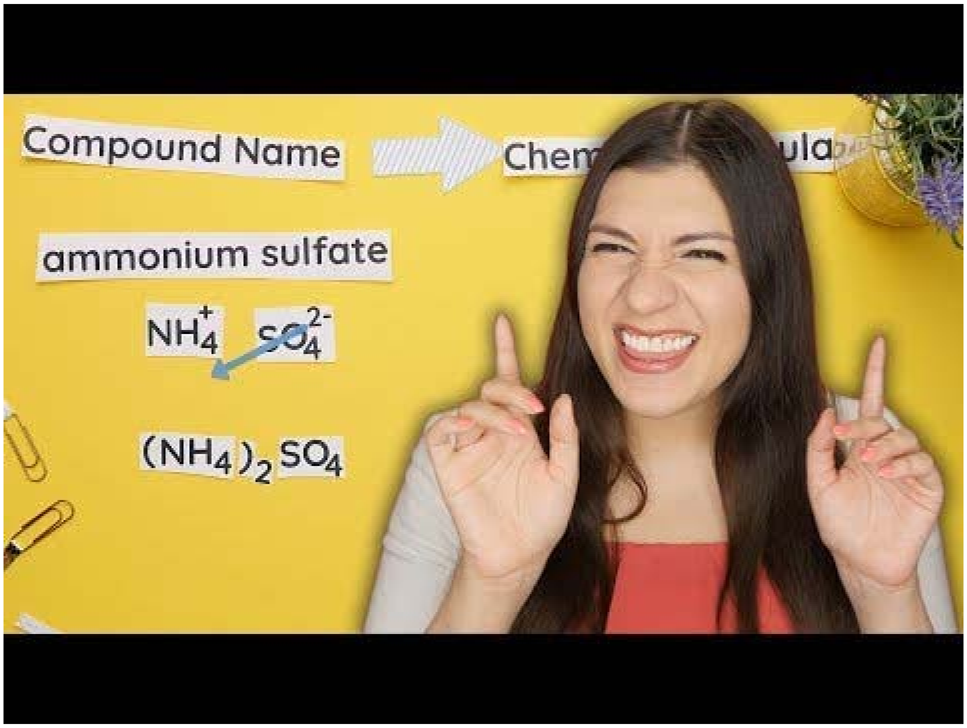
For example: $Na^+ + SO_4^{2-} \rightarrow Na_2SO_4$

But when Ca^{2+} is combined with PO_4^{3-}, the formula of the compound is:

$$Ca^{2+} + PO_4^{3-} \rightarrow Ca_3(PO_4)_2$$

Here is another video to practice writing ionic compounds with polyatomic ions.

https://www.youtube.com/watch?v=PPfLDdIfOVA



Questions

1. Name the formula of the ionic compound formed from $Fe^{3+}$and each anion. Then name each compound.

a) OH^-

b) CO_3^{2-}

c) NO_2^-

d) PO_4^{3-}

e) $C_2H_3O_2^-$

Ans: a) $Fe(OH)_3$, Iron(III) Hydroxide
b) $Fe_2(CO_3)_3$: Iron(III) Carbonate
c) $Fe(NO_2)_3$: Iron(III) Nitrite
d) FePO4, Iron(III) Phosphate
e) $Fe(C_2H_3O_2)_3$, iron(III) Acetate

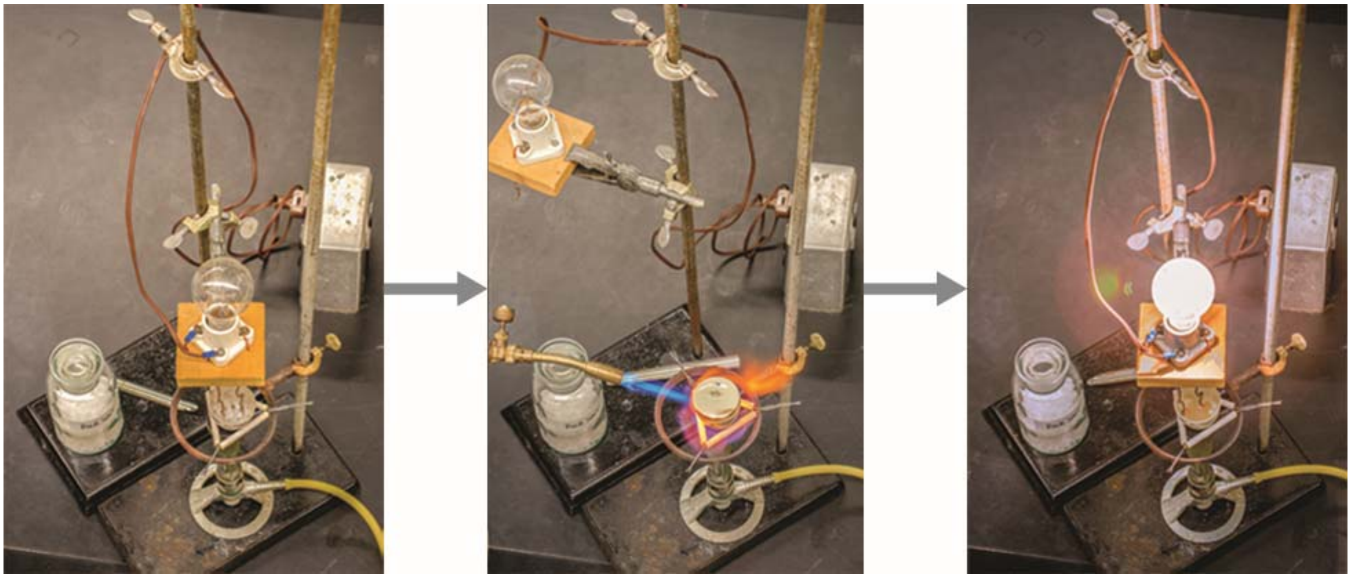
VII. Properties of ionic compounds

Ionic compounds are crystalline solids with the ions arranged to maximize the interactions of the oppositely charged ions. Ionic compounds have high melting and boiling points.

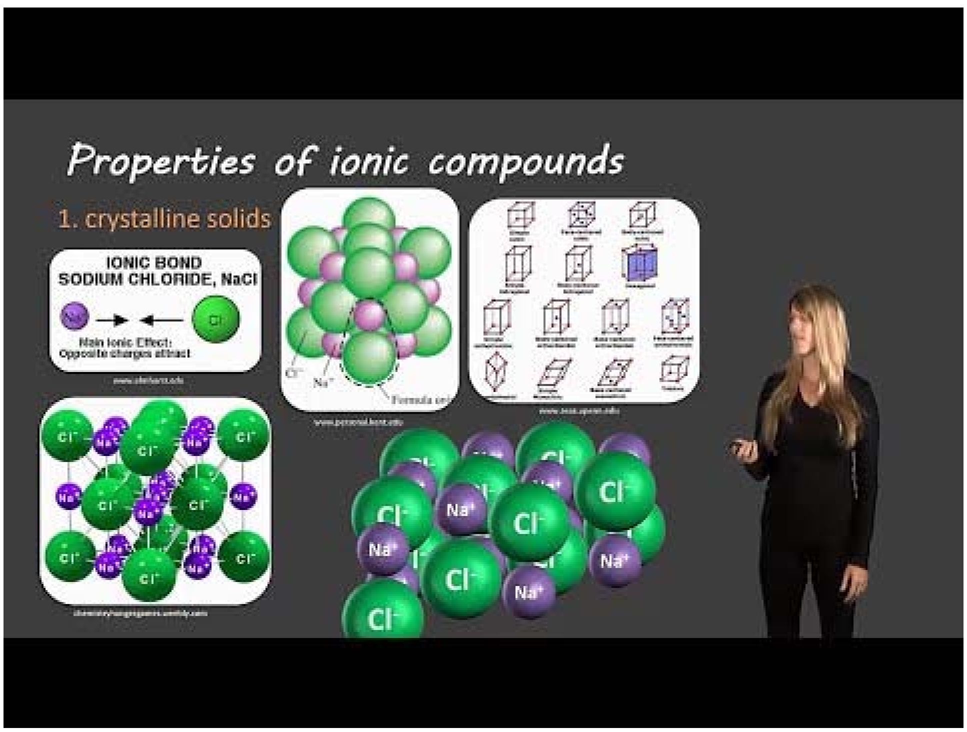
Most ionic compounds are soluble in water and their aqueous solutions conduct electricity. They are called strong electrolyte. Transition metal compounds are colorful, Copper salts are blue and nickel salts are mostly green in color.


$CuSO_4$


$NiCO_3$



Molten sodium chloride conducting electricity.



https://www.youtube.com/watch?v=9FRAtyYK8PI

VIII. Useful Consumer Products and Drugs Composed of Ionic Compounds

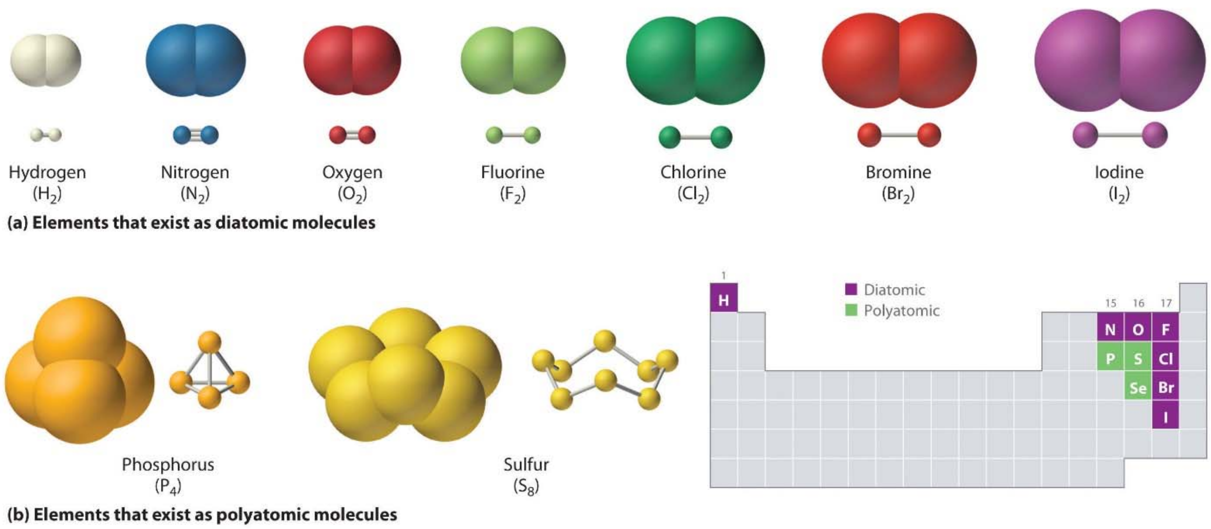
Useful ionic compounds that contain alkali metal cations and halogen anions include KI (iodine supplement), NaF (source of fluoride in toothpaste) and KCl (potassium supplement)
Otwer products contain SnF_2 (fluoride source in toothpaste) Al_2O_3 (Abrasive in toothpaste) and zince oxide (sunblock agent). Useful ionic compounds with polyatomic ions include $CaCO_3$ (antacid and calcium supplement),

Name	**Formula**	**Medical Use**
$AgNO_3$	Silver Nitrate	Antibiotic
$BaSO_4$	Barium Sulfate	Radiopaque material for human consumption
NaI	Sodium Iodide	Thyroid Treatment
Li_2CO_3	Lithium Carbonate	Bipolar disorder treatment
$NaHCO_3$	Sodium Bicarbonate	Antacid
SnF_2	Tin(II) Fluoride	Teeth strengthening agent
ZnO	Zinc Oxide	Sunscreen lotion

IX. Formation of Covalent Compound

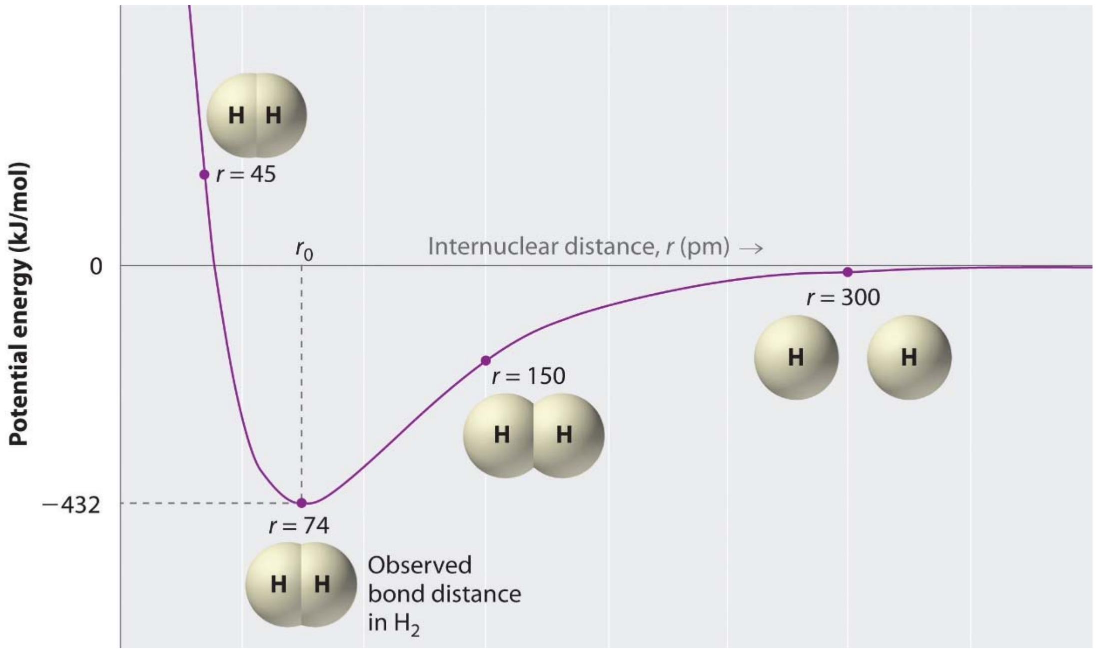
A covalent bond is formed when a valence electron pair is shared between two atoms with a very low difference in electronegativity. As mentioned before, most covalent bonds are formed between two nonmetals or between a nonmetal and a metalloid. When electrons are shared, both atoms gain stability by achieving octet. The shared electron pair is called a bond and the electron pairs involved in creating the bond is called **bonding electrons**. Other valence electrons that don't participate in bond formation are called **lone pair** of electrons.

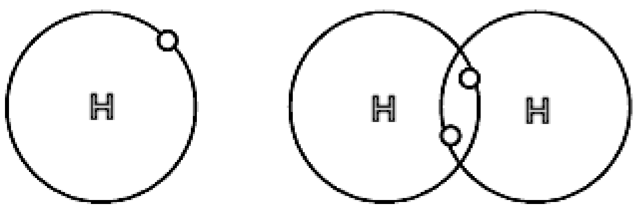
There are some exception to the octet rule, Hydrogen follows duet rule. The number of covalent bond possible in a homonuclear diatomic molecule is always 8-valence electrons. If two or three electron pairs are shared, there are called double and triple bonds respectively.



A single bond is represented by a line "____" and a double bond is represented by double line "====". A lone pair of electrons is indicated by two dots. ".." on the atom.

The following diagram represents potential energy diagram of H_2 molecule formation. Although two isolated nonmetals are electrically neutral, ideally they don't experience any opposite charge attraction. But when the two atoms of same or different nonmetals are brought close together, due to their momentary displacements of protons and electrons, gradually when they reach an optimum distance the covalent single bond is formed. The internuclear distance between two bonded atoms can be determined experimentally and called bond length.

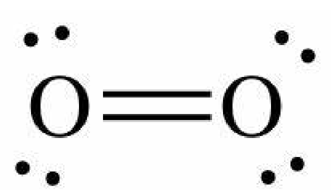


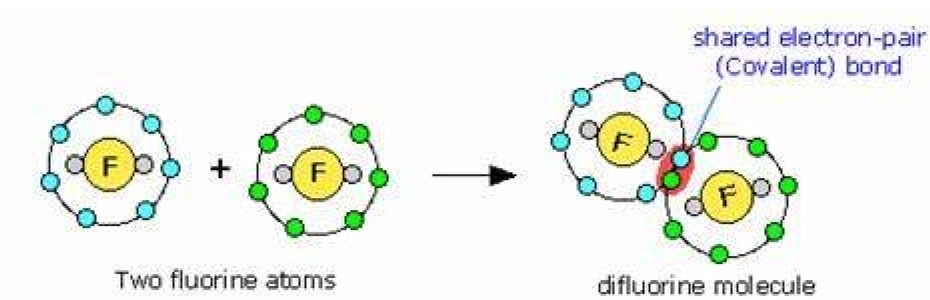


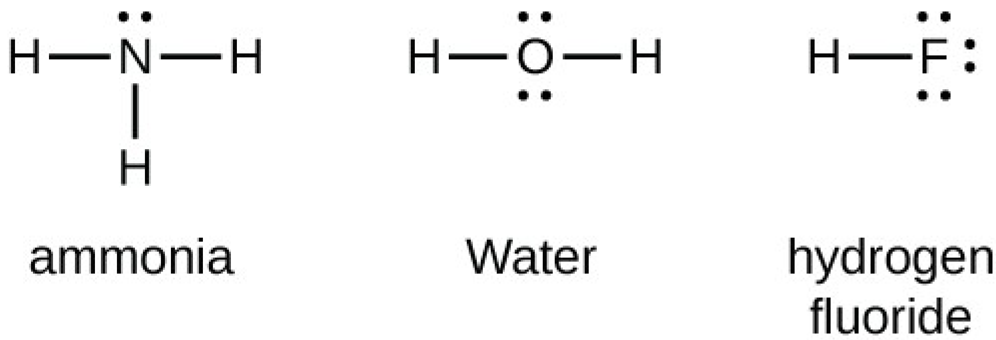
The overlap of above two electrons in H_2 molecule created a single bond. Both H atoms share this pair of electrons.

In the example below, to achieve octet, oxygen molecule must form double bond. Therefore, 4 bonding electrons and 8 lone pair of electrons are present in O_2 molecule.

Below are some examples of covalent molecules with bonding and nonbonding(lone pair) electrons.

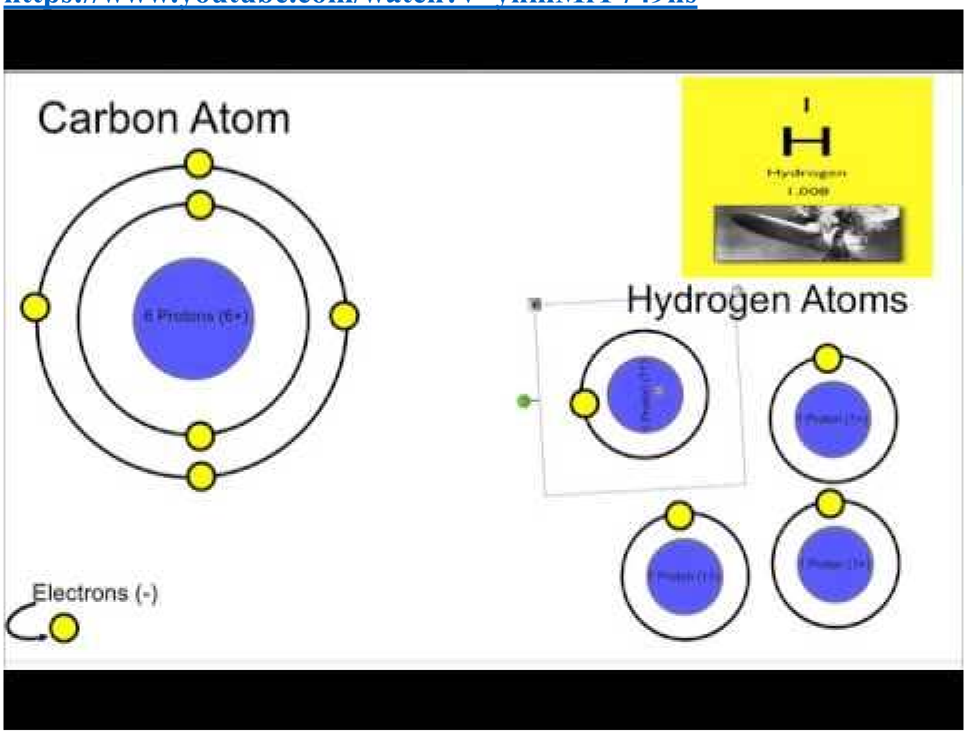






The following video may be helpful in understanding covalent bonding.

https://www.youtube.com/watch?v=yhmMrF749ns



Questions

1. How many covalent bonds are possible for Carbon?
2. How many lone pair and bond pairs of electrons present in N_2 molecule?

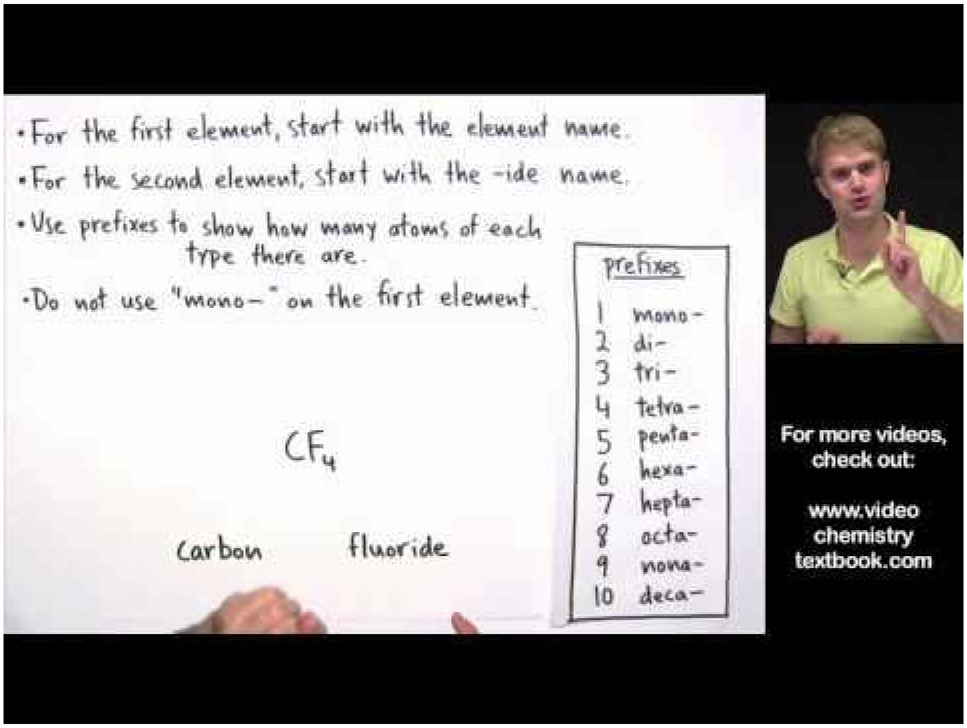
Ans:1. 4   2. Bond pair: 3, lone pair: 2

X. Nomenclature of Binary Covalent Compounds

Prefixes are used to write the name of the covalent molecules. The first nonmetal is usually less electronegative and contains original name, the second nonmetal is written with -ide suffix along with the root name. For example: PH_3 is called as Phosphorus Trihydride and H_2S is called Hydrogen sulfide. Following prefixes are added based on the number of atoms of each nonmetal.

Number of atoms	Prefix
1	Mono (generally not used, exception: CO: carbon monoxide)
2	Di-CO_2
3	Tri-NCl_3
4	Tetra
5	Penta
6	Hexa
7	Hepta
8	Octa
9	Nona
10	Deca

https://www.youtube.com/watch?v=DejkvR4pvRw&t=3s



Questions:

1. Name the following compounds:
   CF_4, NCl_5, P_4O_{10}

2. Write the formula of the following compounds:
   Dichlorine heptoxide

Ans: 1. CF_4: Carbon tetrafluoride, NCl_5: Nitrogen Pentachloride, P_4O_{10}: Tetraphosphorus Decoxide
2. Cl_2O_7

The following activity has been taken from AACT

Simulation: Ionic and Covalent Bonding

Background

In this investigation you will bond select atoms. Based upon the types of atoms that you choose to combine, you will create either an ionic compound or a covalent compound. You will have the opportunity to analyze the differences between these different types of compounds and to predict the number of atoms needed to create each, as well as learn how to appropriately name them.

1. Describe the difference between an atom and a molecule:

2. Where are metal atoms located on the periodic table? Where are non-metal atoms located on the periodic table?

3. What subatomic particle(s) participate in chemical bonding?

4. In your own words, define *valence electron*.

5. How can you determine the number of valence electrons in an atom using the periodic table?

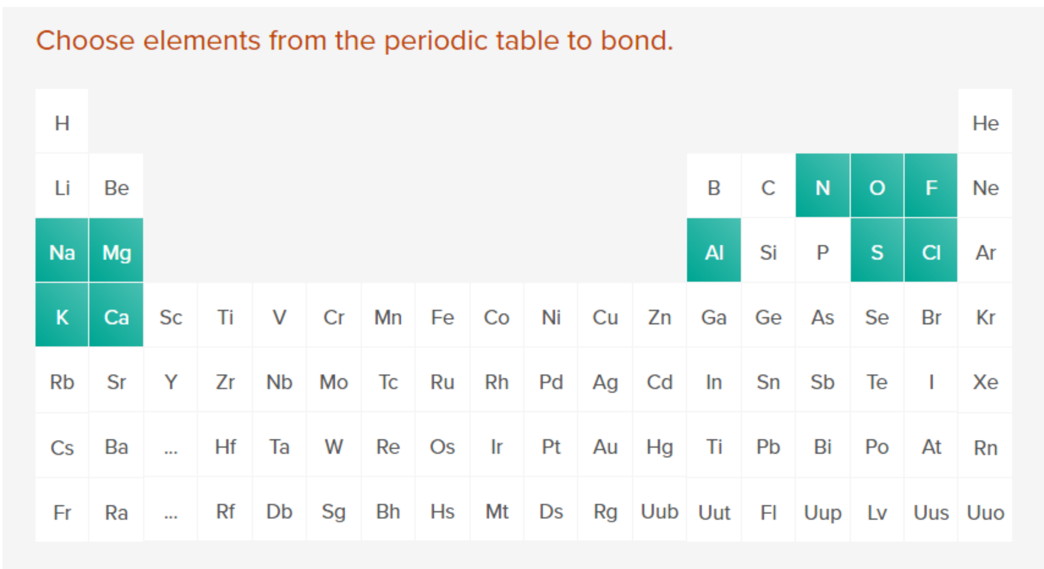
6. Draw a Lewis Dot Structure for the following atoms:

a. Strontium (Sr)
b. Carbon (C)
c. Iodine (I)
d. Xenon (Xe)

*Check your answers before moving on to the next portion of the activity.

Procedure

Using your computer, tablet or mobile device, navigate to the website: http://www.teachchemistry.org/bonding. You should see the picture below on your screen.



Part 1: Ionic Bonding

1. Choose Sodium (Na).
    a. What *type of element* is it?

    b. How many valence electrons does it have?

2. Choose Fluorine (F).

a. What *type of element* is it?

b. How many valence electrons does it have?

3. Answer the question on the screen, “What type of bond is this combination likely to form?”

   a. Circle: Ionic or Covalent?

   b. Choose the appropriate number of atoms to make the bond. Record the number of each atom below:

4. Watch the final animation closely (it will play continuously).
   a. Describe the change in the number of valence electrons in the atoms as the bond is successfully formed:

   b. What does the positive (+) charge indicate (mention specific subatomic particles in your answer)?

   c. What does the negative (-) charge indicate (mention specific subatomic particles in your answer)?

   d. What is the final overall charge?

   e. Record the name and molecular formula for the compound below:

Reset the selected data using the reset symbol.

5. Choose Calcium (Ca).
   a. What *type of element* is it?

   b. How many valence electrons does it have?

6. Choose Chlorine (Cl).
    a. What *type of element* is it?

    b. How many valence electrons does it have?

7. Answer the question on the screen, "What type of bond is this combination likely to form?"

    a. Circle: Ionic or Covalent?

    b. Choose the appropriate number of atoms to make the bond. Record the number of each atom below:

8. Watch the final animation closely (it will play continuously).
    a. Why were more than 2 total atoms needed to create this compound?

    b. Explain what happened to the valence electrons in <u>each</u> atom.

    c. What is the final overall charge?

    d. Record the name and molecular formula for the compound below:

    e. Have you noticed a pattern between the charge of the ion and the number of valence electrons it has? Explain how you can predict the charge based on the number of valence electrons, or the location of the element on the periodic table.

Reset the selected data using the reset symbol.

9. Using a periodic table, complete the table below, then use the simulation to check each of your predictions:

Atom #1	Number of Valence Electrons	Prediction of charge	Atom #2	Number of Valence Electrons	Prediction of charge	Molecular Formula	Name of compound
Na			**O**				
K			**F**				
Mg			**Cl**				
Ca			**N**				
Al			**S**				

Part 2: Covalent Bonding

1. You will first investigate 5 *diatomic* molecules. Diatomic molecules are made up of 2 atoms.

    a. Select 2 fluorine atoms. How many valence electrons are in each fluorine atom?

    b. Is a fluorine atom a metal or a non-metal?

    c. Did the combination of these atoms create a covalent or ionic bond?

d. How are the valence electrons organized to form a bond between these atoms?

e. How is this different from the ionic bonds formed in the previous part of the activity?

f. What shape does this molecule form?

2.

a. Select 2 oxygen atoms. How many valence electrons are in each oxygen atom?

b. Is an oxygen atom a metal or a non-metal?

c. Did the combination of these atoms create a covalent or ionic bond?

d. How are the valence electrons organized to form a bond between the atoms?

e. How is this bond different from the bond in the fluorine molecule in question 1?

f. What shape does this molecule form?

3. Make predictions in the following table. Once completed, check your answers using the simulation.

Lewis dot structure for single atom	Cl	S	N
Lewis dot structure for diatomic molecule			

(Cl_2, S_2, N_2)			
Molecular formula			
Name of shape			

4. More than two atoms can also be combined to form a covalent molecule. These molecules may form different shapes and will also follow a particular naming system. Select the following combinations of atoms, and complete the rest of the table as you interact with the simulation:

1^{st} atom choice	2^{nd} atom choice	Predict Formula	Molecular Name	Shape
S	F			
N	Cl			
Cl	F			

Part 3: Critical thinking

1. What are the differences between ionic and covalent bonds? Be sure to refer to *valence electrons* in your response.

2. How is naming ionic and covalent compounds different? Use specific examples in your answer.

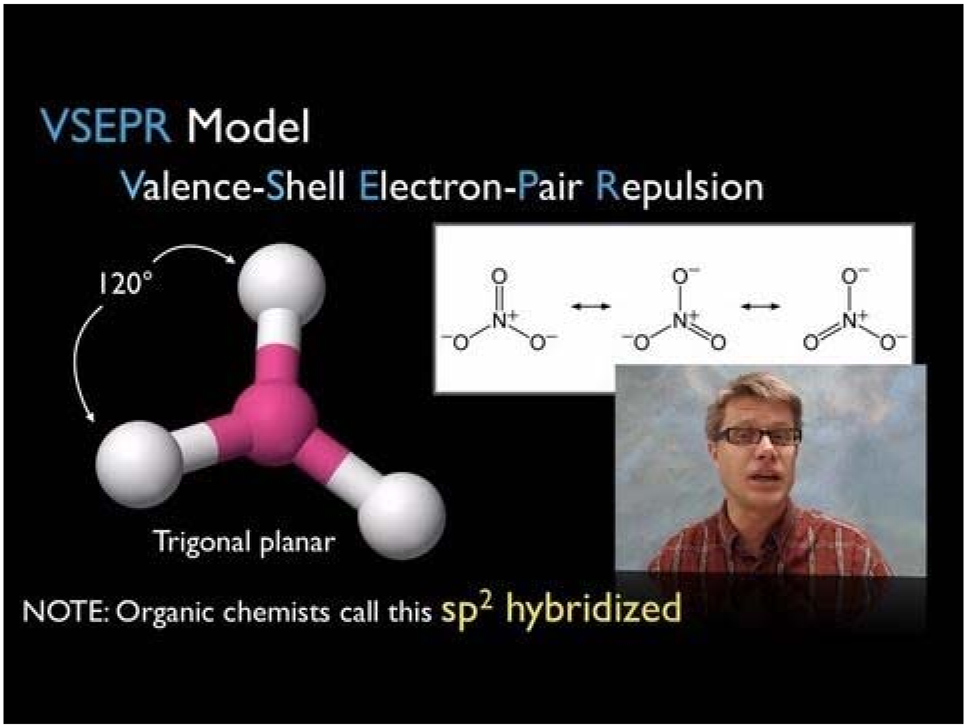
3. Based on your knowledge of ionic and covalent bonds, complete the missing portions of the following table:

Name	Formula	Ionic or Covalent?
Beryllium bromide		
	PF_3	
Sulfur diiodide		
Strontium Phosphide		
	Cs_3N	
	H_2O	

XI. Lewis Structures

Lewis Structures are electron dot representations of molecules. Two-electron bonds are drawn with a solid line and non-bonded electrons are drawn with dots.
Lewis structures contain only valence electrons. H gets two electrons and main group elements generally get eight electrons.
After placing all electrons in bonds and lone-pairs, it may be necessary to use lone pairs from terminal to form multiple bonds if an atom does not have an octet.

https://www.youtube.com/watch?v=xNYiB_2u8J4



Stepwise directions for writing Lewis Dot structures:

1. Determine the total number of valence electrons
2. Determine which atom is the central atom (usually less electronegative atom is central atom. Hydrogen and fluorine can never be central atom)
3. Connect all other atoms to the central by a single covalent bond
4. Satisfy all the surrounding atoms with eight electrons by adding lone pair
5. Verify the total number of valence electrons of the structure. Extra electron should go to the central atom as lone pair.
6. If the central atom doesn't have octet then share the lone pair from terminal atoms with central by making multiple bonds.
7. In case of ions, add or subtract valence electrons based on the charge of the ion.

Below are some examples of drawing Lewis dot structures.

Example#1: CCl_4

1. Total number of valence electrons: C → 4 + Cl → 7*4=28+ 4= total 32 electrons.

2. Since C is less electronegative, it should be the central atom.

3. Then after connecting all four Chlorine atoms by a single bond, the number of electrons remaining= 32-(4*2)= 32-8= 24 electrons.

4. Since each chlorine has only two electrons (a single bond) surrounding it, they require 6 more electrons as lone pair i.e. 3 pairs of lone pairs should be placed on each chlorine. Four chlorine atoms should have 6*4= total 24 remaining electrons.

5. After satisfying octet for the surrounding atom, there is no extra electron to put on central atom.

6. Central atom C has achieved octet through 4 bonds (4*2= 8)

7. There is no need to form multiple bond.

Example #2: CO_2

1. Total number of valence electrons: C → 4 + ClO→ 6*2=12+ 4= total 16 electrons.

2. Since C is less electronegative, it should be the central atom.

3. Then after connecting two Oxygen atoms by a single bond, the number of electrons remaining= 16-(2*2)= 16-4= 12 electrons.

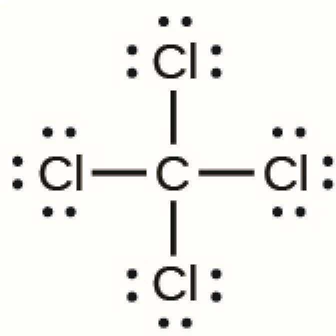
4. Since each oxygen has only two electrons (a single bond) surrounding it, they require 6 more electrons as lone pair i.e. 3 pairs of lone pairs should be placed on each chlorine. Two oxygen atoms should have 6*2= total 12 remaining electrons.

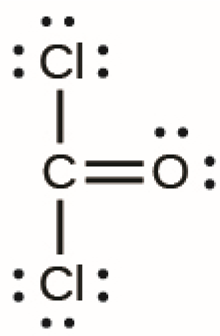
5. After satisfying octet for the surrounding atom, there is no extra electron to put on central atom.

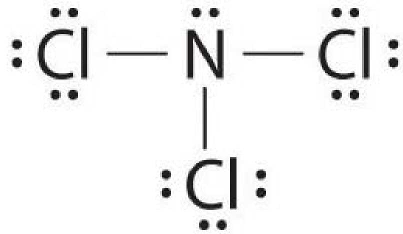
6. Central atom C has **not** achieved since there are only four electrons surrounding carbon by 2 bonds. (2*2= 4)

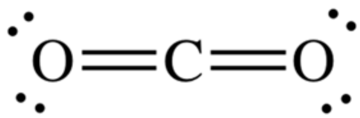
7. There is a need to form multiple bond. Each oxygen's one lone pair is shared with carbon by making a bond between carbon and oxygen. This way Carbon gets extra 4 electrons to achieve octet.

Cn you try other two examples: $COCl_2$ and NCl_3?



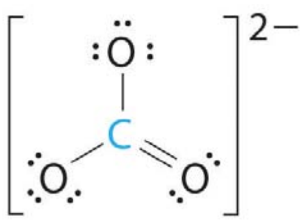






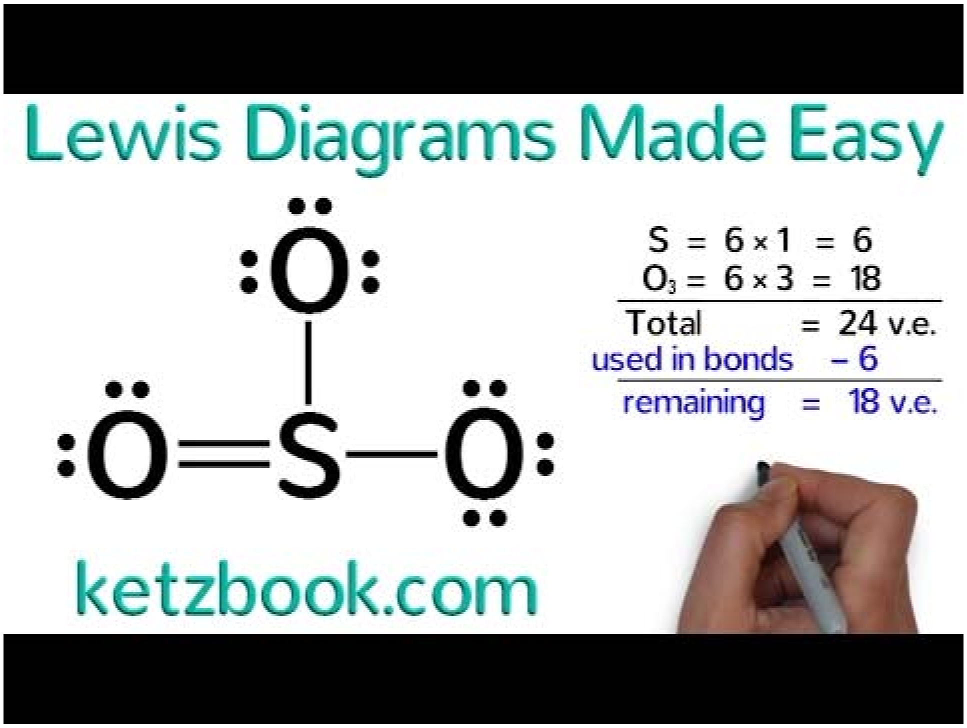
In the following example: CO_3^{2-} ion has total= 4*1 + 6*3 +2= 24 electrons.

Carbon is the central atom, after connecting all other atoms and satisfy the terminal atoms with 8 electrons, there is no more extra electron left to put on central. But Carbon still doesn't have octet. So to satisfy the carbon with 8 electrons any terminal oxygen atom's lone pairs are shared to form a double bond.
Normally ions are written with bracket with the charge mentioned outside the species.



Lewis structure

https://www.youtube.com/watch?v=cIuXl7o6mAw



In some cases exception to octet rule is found. For example, Hydrogen follow duet rule, B oron requires six electrons. Any period three nonmetal (P, S, Cl) and higher elements can hold more than eight electrons by expanding their partially filled or empty “d” orbitals.
https://www.youtube.com/watch?v=Dkj-SMBLQzM

Questions:

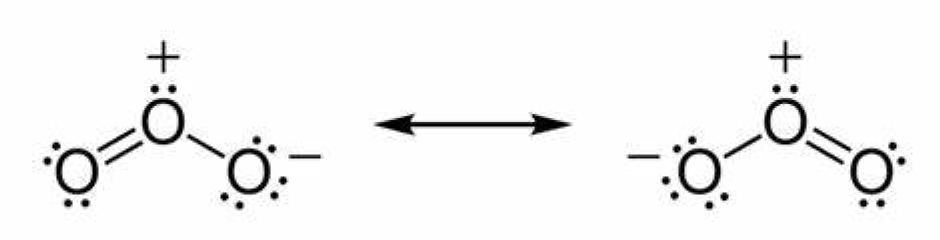
1. Draw the Lewis dot structures of following molecules

a) SO_2
b) O_3
c) SO_3
d) PCl_3
e) H_2CO
f) C_2H_4
g) CH_3OH

Ans: See your instructor

XII. Resonance structure

Resonance structures are multiple Lewis dot structures having the same arrangement of atoms but a different arrangement of electrons. This is possible in case of multiple valid Lewis dot structures.The hybrid is a composite of all resonance structures that spreads out electron pairs in multiple bonds and lone pairs. In resonating structures, electron pairs move from one part of the atom to the other part. Resonating structures are shown by a double headed arrow.



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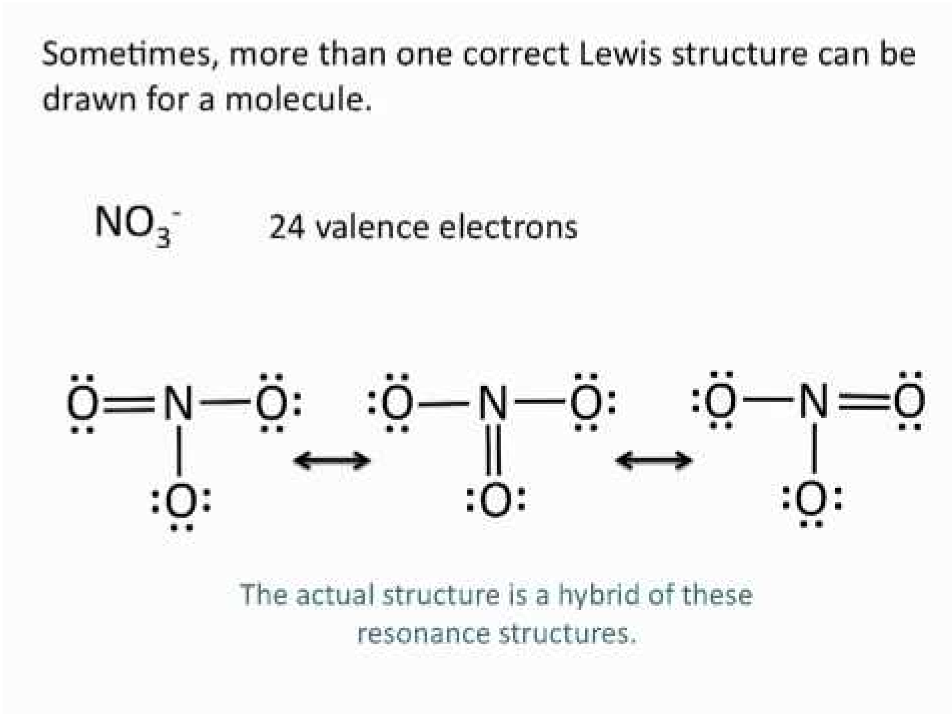


All contributing structures must:
1. Have the same number of valence electrons.
2. Obey the rules of covalent bonding.
    - No more than 2 electrons in the valence shell of H.
    - No more than 8 electrons in the valence shell of a 2nd period element.
    - 3rd period elements, such as P and S, may have up to 12 electrons in their valence shells.
3. Differ only in distribution of valence electrons; the position of all nuclei must be the same.
4. Have the same number of paired and unpaired electrons.

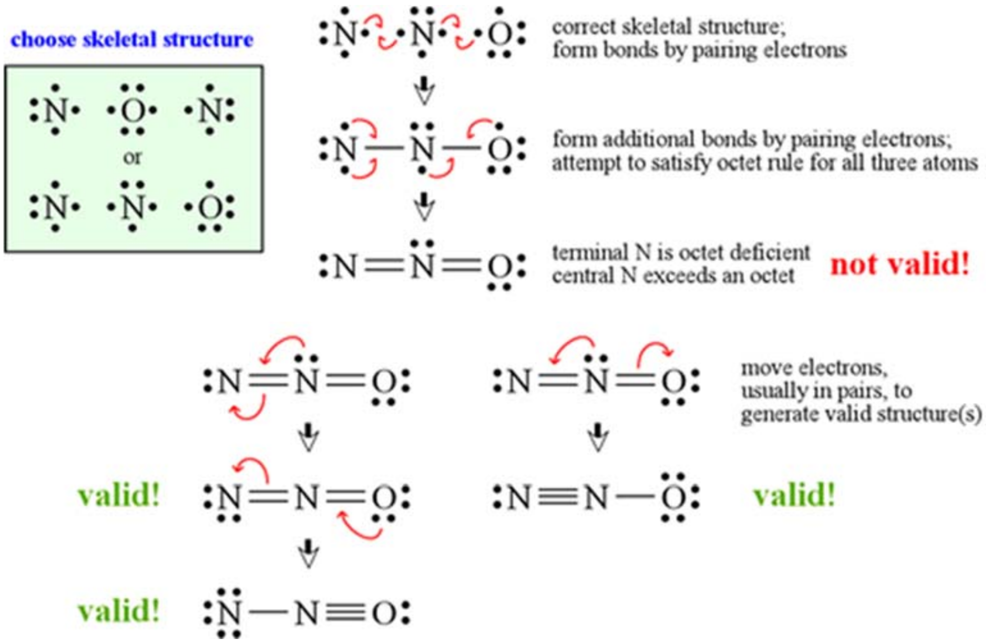
Curved arrow: A symbol used to show the redistribution of valence electrons:

- the tail of the arrow identifies a pair of electrons whose location is changing.
- the head of the arrow identifies the the new location of the involved pair of electrons.

In using curved arrows, there are only two allowed types of electron redistribution:

- from a bond to an adjacent atom.
- from an atom to an adjacent bond.

Here is another example of resonance.



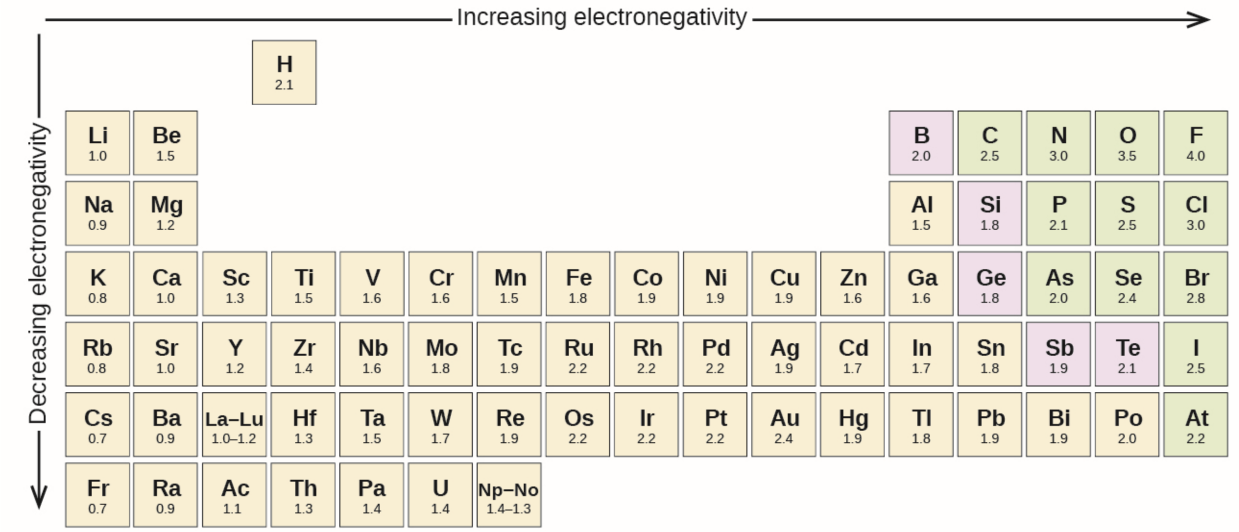
Questions:

1. Draw resonance structures of
   a) NO_2-, C_6H_6

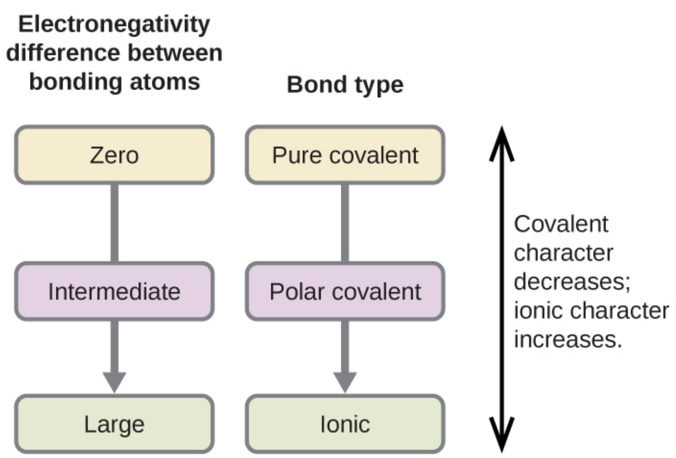
Ans: see your instructor

XIII. Electronegativity & Bond Polarity

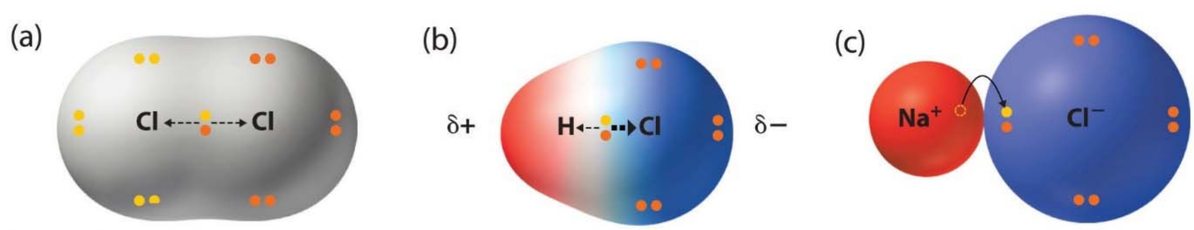
Electronegativity is a measure of an atom's attraction for electrons in a bond. Electronegativity is defined as the ability of an atom to attract shared pair of electrons in a bon towards itself. Metals have tendency to lose electrons, therefore low electronegativity, nonmetals have tendency to gain electrons, and high electronegativity. In general, electronegativity increases from electro right across the periodic table and decreases from top to bottom with increasing metallic character. We assign small charge δ- to the more electronegative atom and δ+ to the less electronegative atom.
Scientist Pauling first determined the absolute values of electronegativity



When two atoms have same electronegativity value or the difference is less t han 0.5 units, the electrons are equally shared and bonds are nonpolar. when two atoms have very different electronegativity values- a difference of 0.5-1.9 units; the electrons are unequally shared and the bond is polar.



Bond Type	Electronegativity Difference
pure covalent	< 0.4
polar covalent	between 0.4 and 1.8
ionic	> 1.8

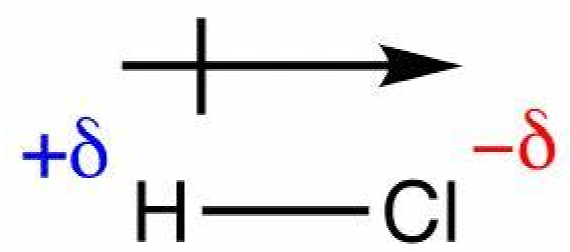


Nonpolar covalent bond
Bonding electrons shared equally between two atoms.
No charges on atoms.

Polar covalent bond
Bonding electrons shared unequally between two atoms.
Partial charges on atoms.

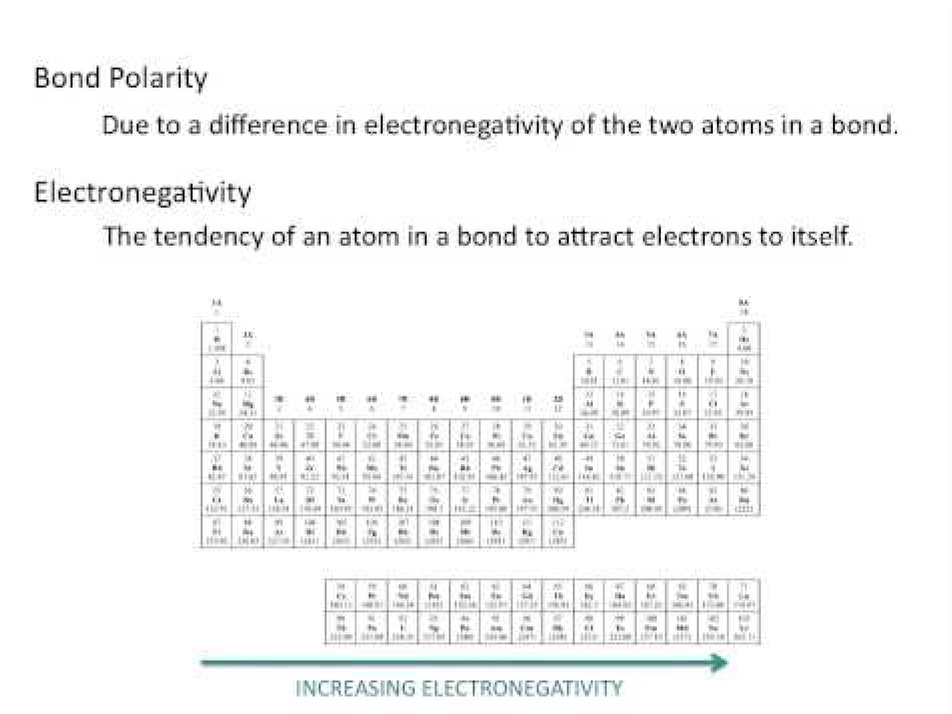
Ionic bond
Complete transfer of one or more valence electrons.
Full charges on resulting ions.

A polar bond is written with partial charge on both sides of the atoms forming the bond. Another way to show the bond polarity is b drawing dipole i.e. an arrow point to the more electronegative atom displayed on top of the bonded atoms with a vertical line on the other end.



Watch the following video:

https://www.youtube.com/watch?v=_T9iZWhZTY8&t=2s



Questions

1. Classify each bond as nonpolar covalent, polar covalent or ionic

A) N-Li
B) P-H
C) N-O
D) Cl-Cl

2. Indicate the dipole and partial charges in following bonds
A) C-Mg
B) I-Cl

Ans: 1. A) Ionic, B) Polar covalent C) Polar Covalent, D) Covalent
2.A) C(-)___Mg(+) , B) I(+) __Cl(-)

XIV. Molecular Shape and VSEPR

This theory helps to determine the shape of polyatomic covalent molecules. In a polyatomic molecule, central atom solely determines the shape. The total number of bond pair and lone pair determines the geometry of the molecule. Shape is determined based on the number of lone pair on central atom.

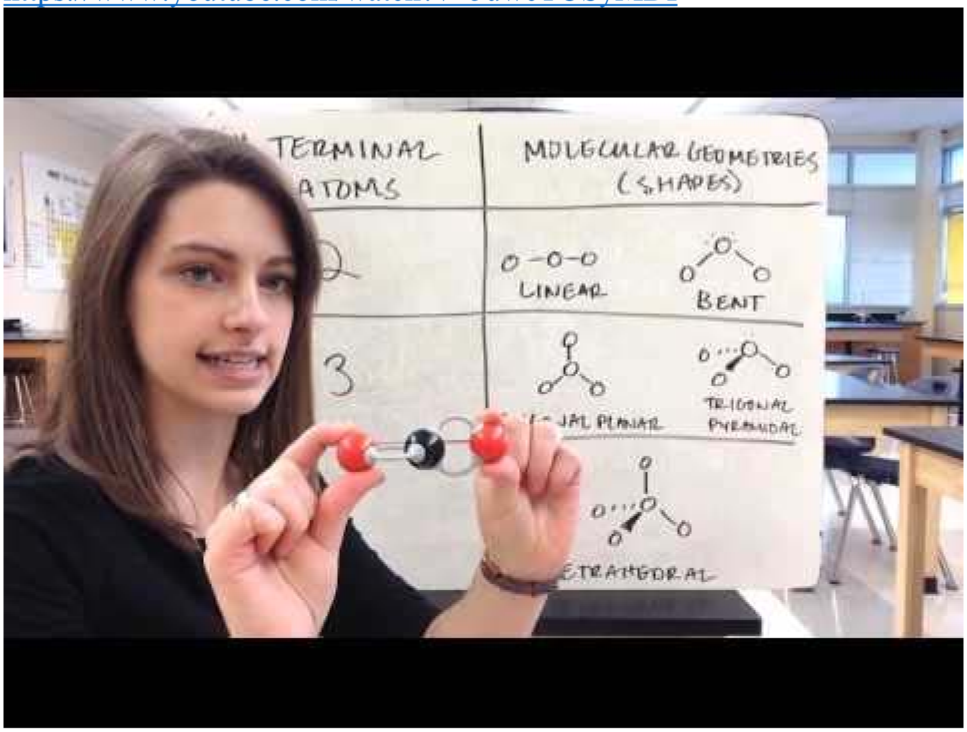
The complete form of VSEPR is Valence Shell Electron Pair Repulsion Theory. There are three postulates of VSEPR theory

1. Electrons are arranged in polyatomic molecule I such a way that the repulsion between atoms is minimized.
2. Any Bond pair or lone pair is considered as one density area surrounding area of electron density. VSEPR theory doesn't distinguish between single and multiple bond.
3. Lone pair on central atom is considered as electron cloud and repels the bond pair of electrons causing the bond angle to decrease. The degree of repulsion
Lone pair-lone pair> lone pair-bond pair> bond pair-bond pair

The geometry based on the total number of electron groups surrounding central atom is called electronic geometry. 3D molecular shape changes from electronic geometry when there is(are) lone pair of electrons on the central atom. Bond angle is determined by connecting the two adjacent atoms with a central atom.

To determine the shape, count groups, atoms and lone pairs, and keep the groups as far away from each other as possible.
Two groups- linear, 180 degree, three groups- trigonal planar, 120 degree, four groups- Tetrahedral 109.5 degree.

https://www.youtube.com/watch?v=5uwcTOSyMDI



Here is the complete VSEPR table with 2, 3, 4, 5 and 6 regions of electron density area surrounding the central atom.

Here is the complete VSEPR chart.

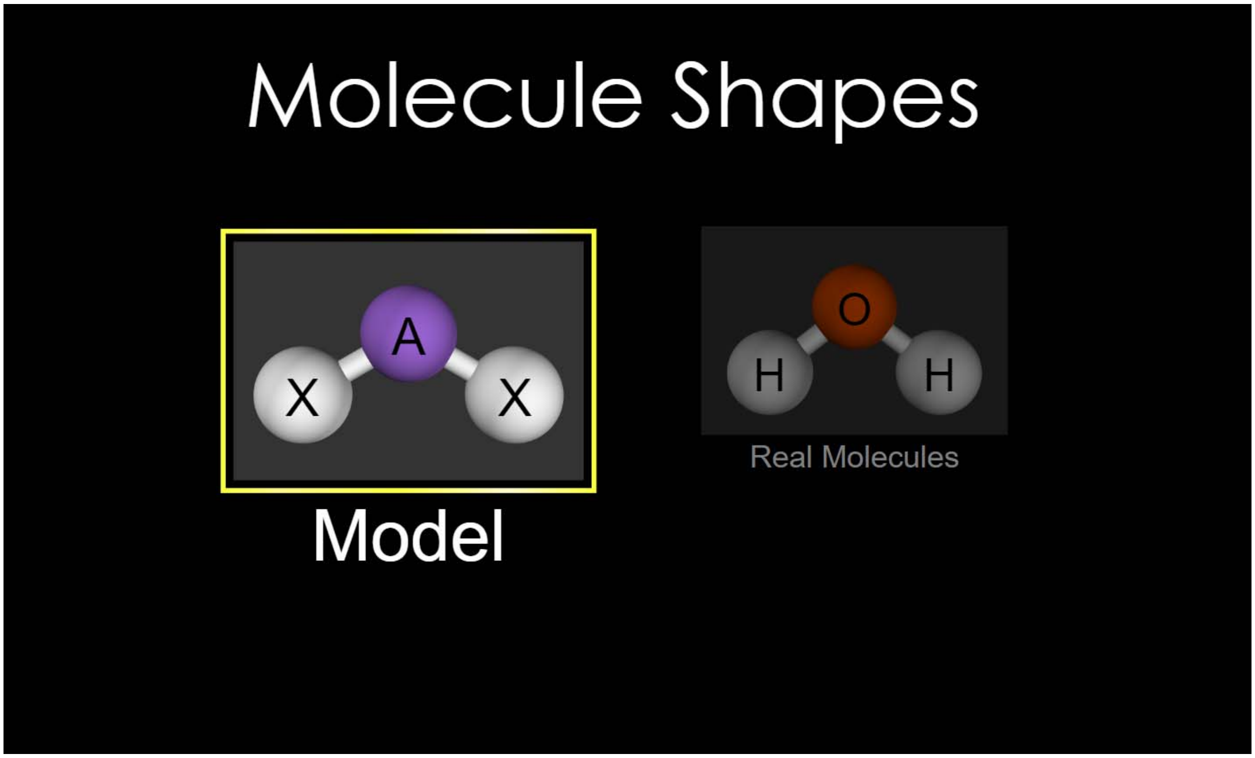
VSEPR Theory – 3D shape of the polyatomic molecules

# Areas of e- density or steric number			Electronic Geometry	Molecular Shape	Bond Angles
Total	**Bond pairs**	**Lone pairs**			
2	2	0	Linear	Linear	180°
3	3	0	Trigonal planar	Trigonal planar	120º
3	2	1	Trigonal planar	Angular	Less than 120°
4	4	0	Tetrahedral	Tetrahedral	109.5°
4	3	1	Tetrahedral	Trigonal pyramidal	Less than 109.5°
4	2	2	Tetrahedral	Angular	Less than 109.5°

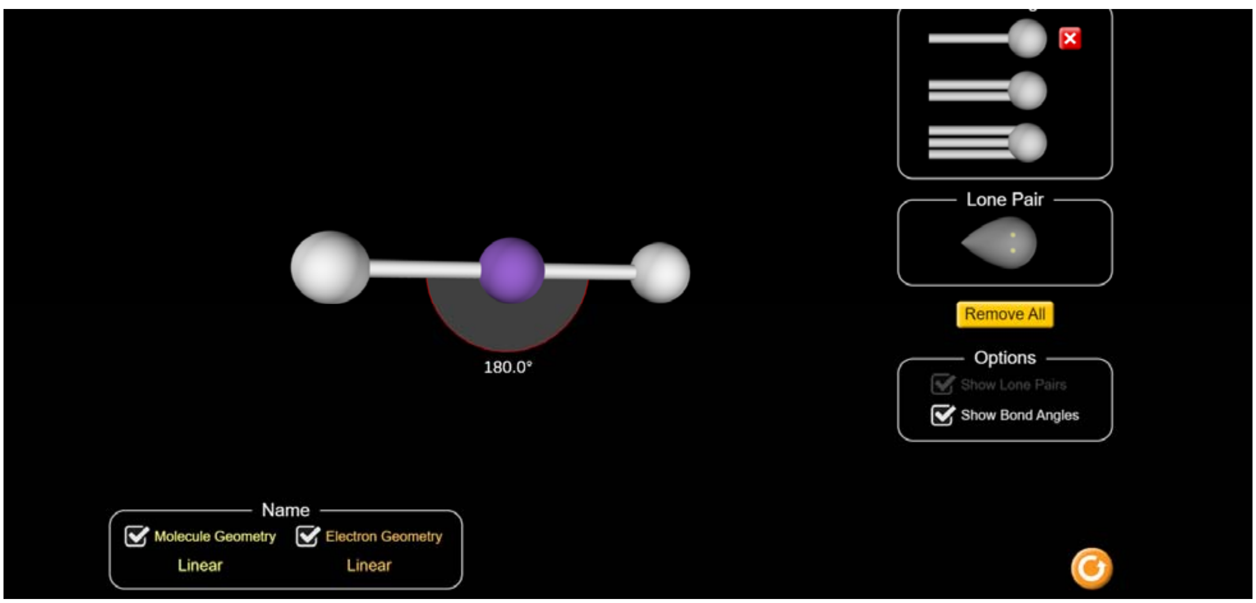
We will try to understand VSEPR thorugh the simulation below:

Description: Open the Phet simulation link to VSEPR activity
https://phet.colorado.edu/sims/html/molecule-shapes/latest/molecule-shapes_en.html

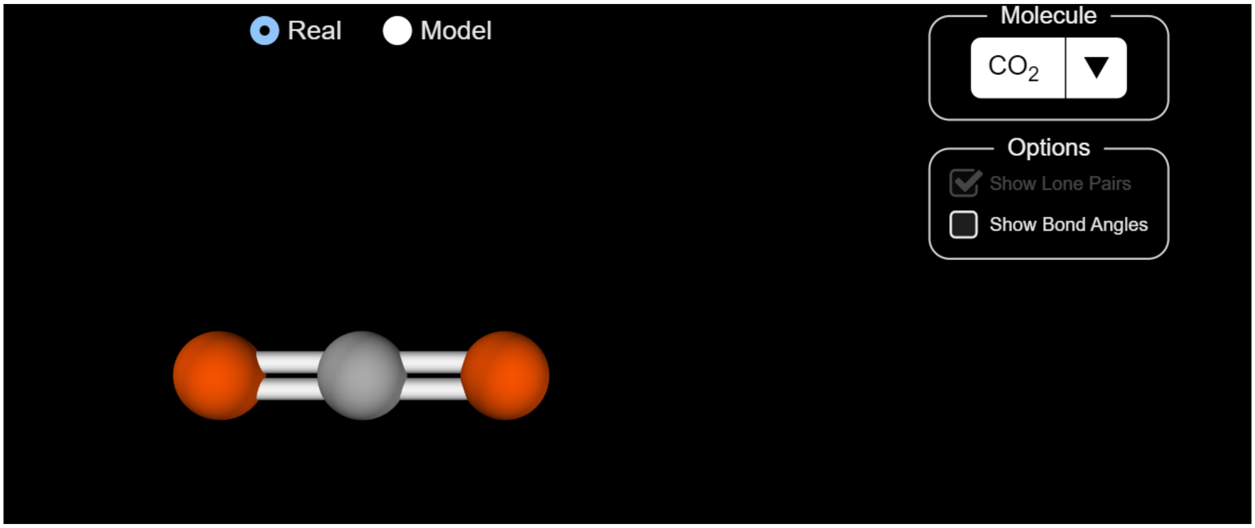
[Advanced <iframe src="https://phet.colorado.edu/sims/html/molecule-shapes/latest/molecule-shapes_en.html" width="100%" height="600"]



1) Two electron groups:
   Click on the "model " and observe the two electron groups electron geometry surrounding the central atom. Click on the elctron geometry, molecular geometry and bond angle.

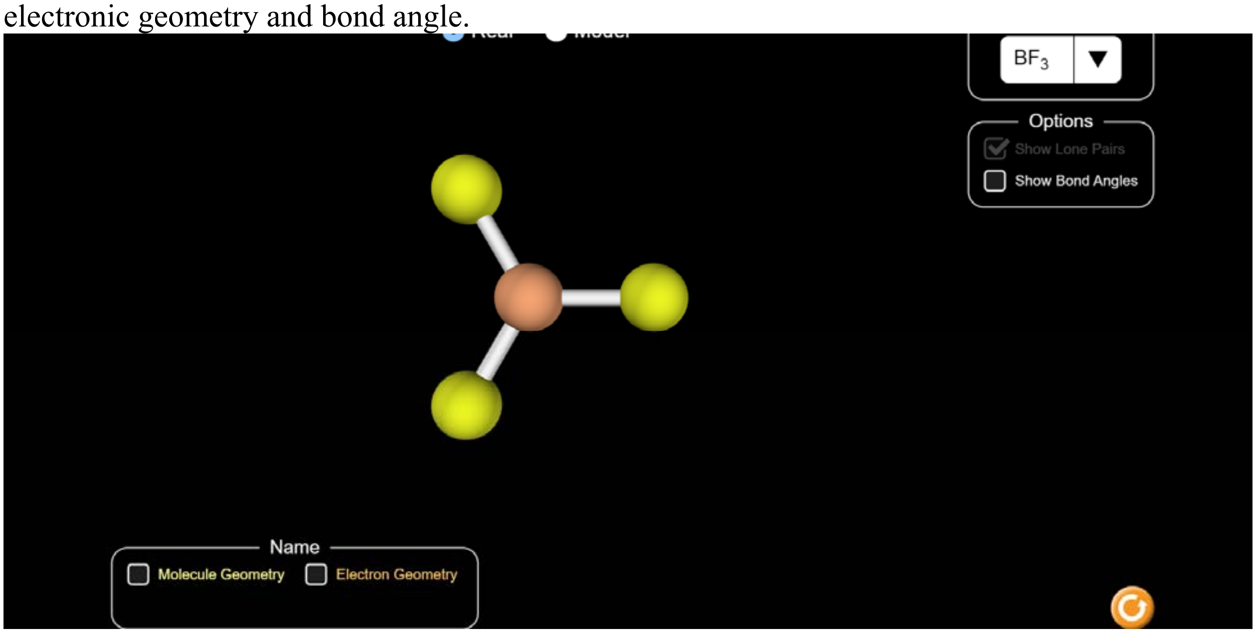


Then go to the real molecule section and choose CO_2. Can you predict the molecular, electronic geometry and bond angle?

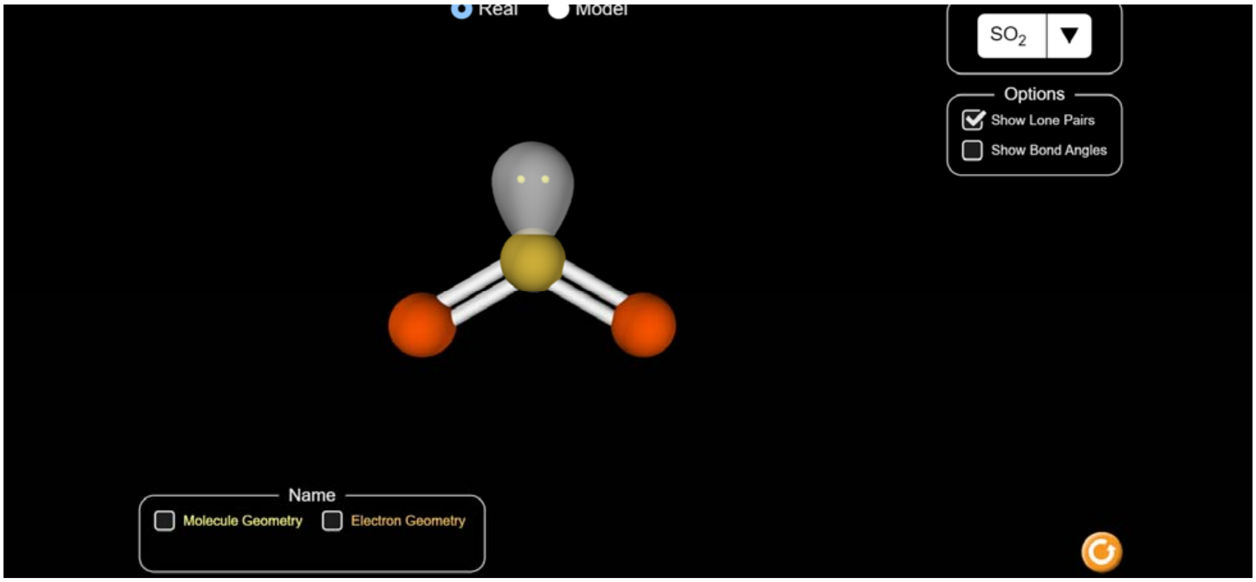


2) Three electron groups:

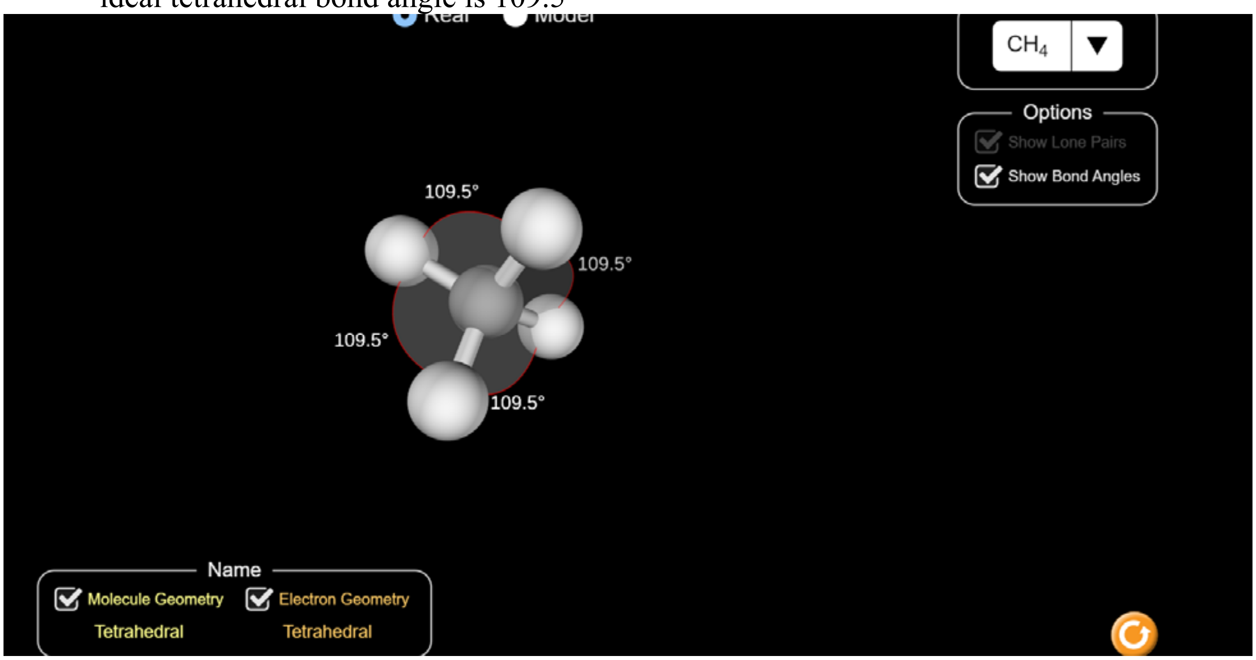
Choose BF3 molecule under real molecule section and click on the molecular geometry, electronic geometry and bond angle.



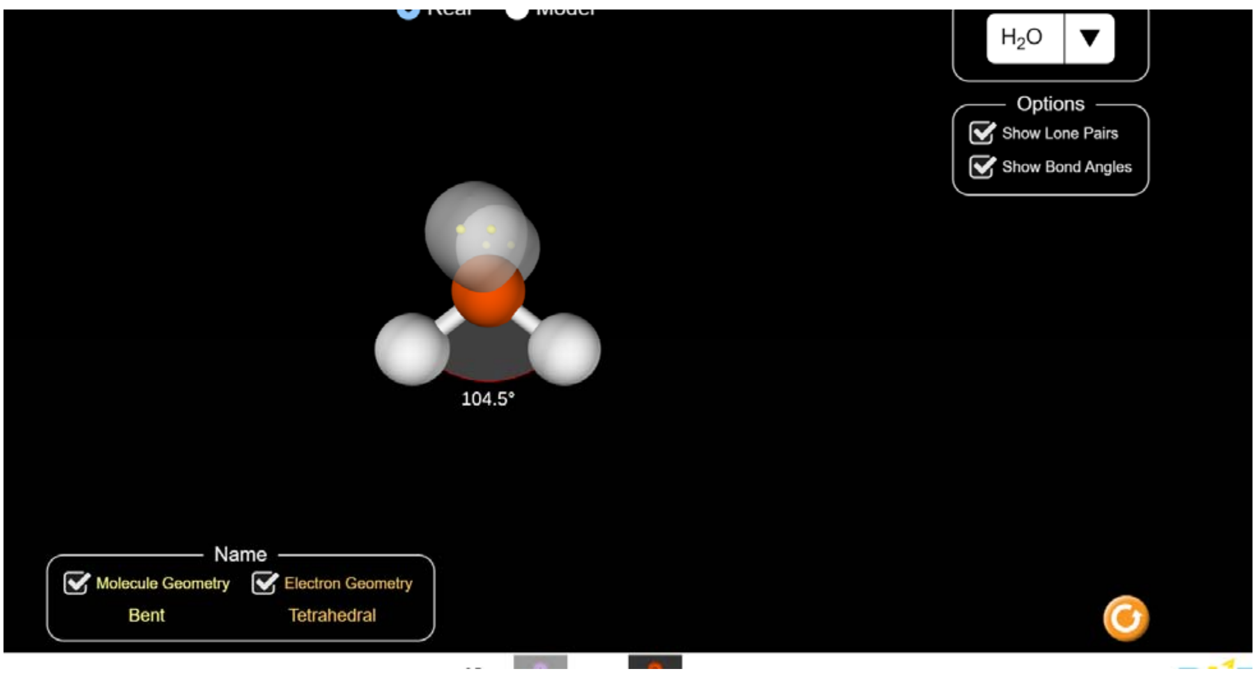
Now choose SO_2, in SO_2, total three electron density areas but two of them are bonding and one is lone pair. What do you observe for electronic, molecular geometry and bond angle? Why?



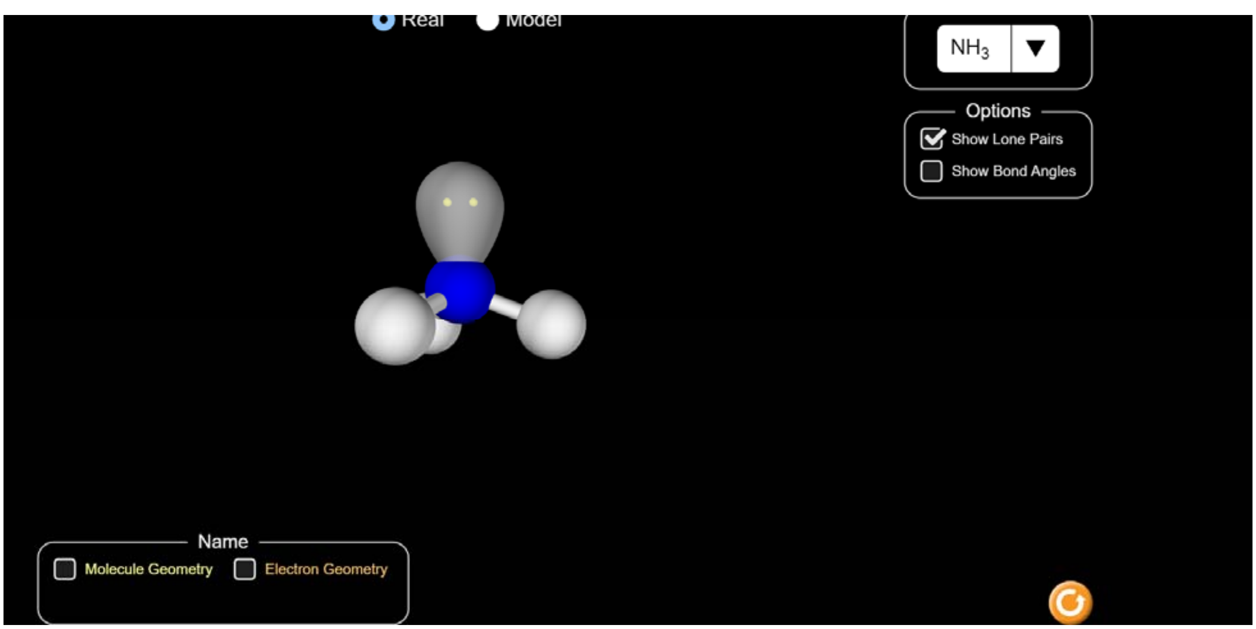
3) Four electron groups:
   In methane, there are total bond groups an no lone pair. For four electron density areas, the ideal geometry is tetrahedral and molecular shape is same as electronic geometry. The ideal tetrahedral bond angle is 109.5^0



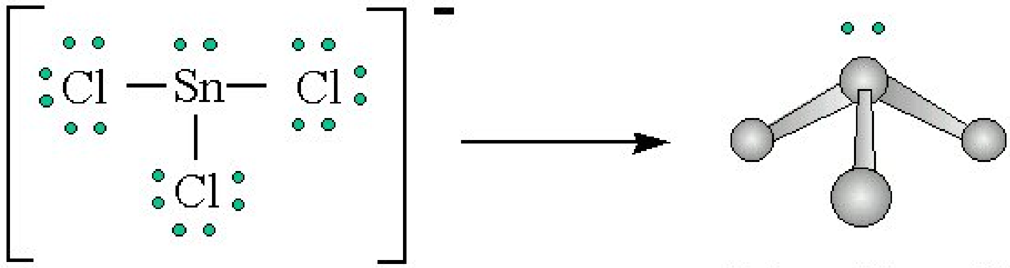
Take the water molecule. H_2O has two bond pairs and two lone pairs, total four electron density groups. The shape of the molecule is bent although the geometry is tetrahedral. The bond angle is 104.5^0 which is less than ideal for tetrahedral geometry (109.5^0) due to presence of two lone lone pairs.



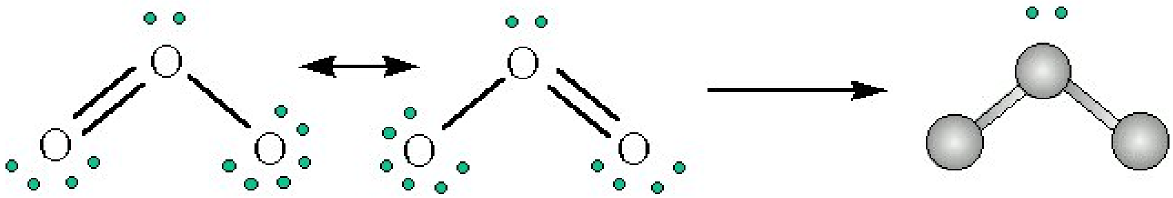
Choose the molecule NH_3. Can you predict the electronic geometry, molecular geometry and bond angle looking at the structure?



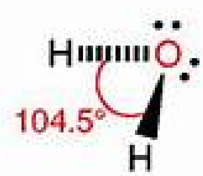
Here are some more examples:



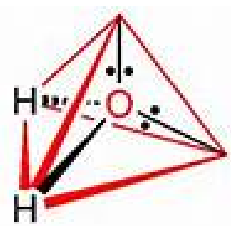
Trigonal Pyramidal



Bent

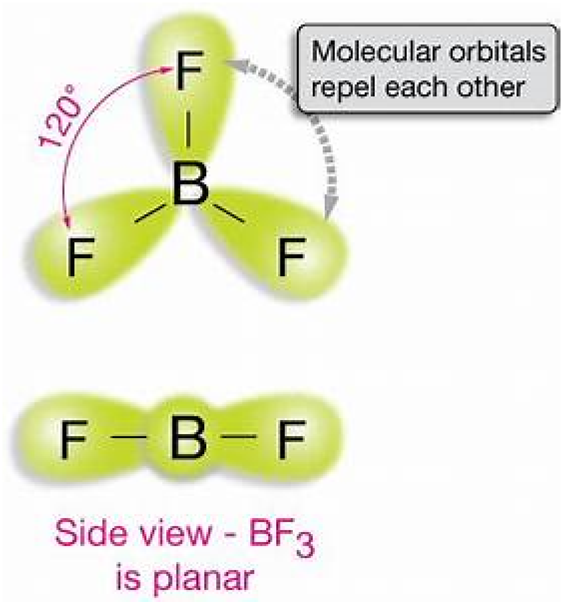


The H–O–H bond angle is 104.5°

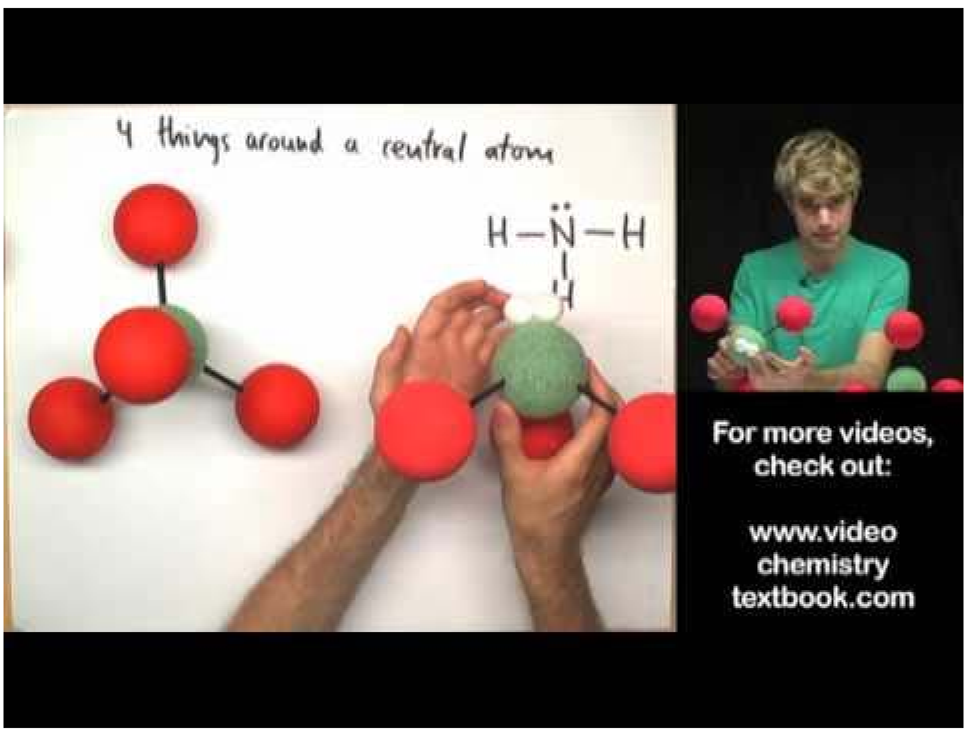


The red lines outline a tetrahedron
Black lines show the electron pairs

The molecular shape of H_2O is bent.



https://www.youtube.com/watch?v=nxebQZUVvTg&t=947s



To practice more problems, watch the video.
https://www.youtube.com/watch?v=xwgid9YuH58

Questions:

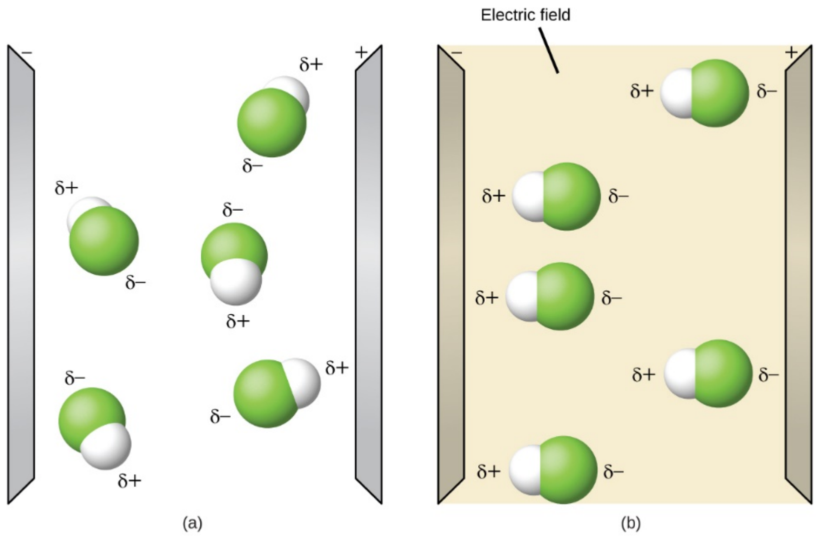
1. For each of the following draw the Lewis dot structure and determine electronic geometry, molecular shape and bond angle.

a) NF_3
b) BI_3
c) CH_4
d) SO_2
e) C_2H_4
f) CO_3^{2-}

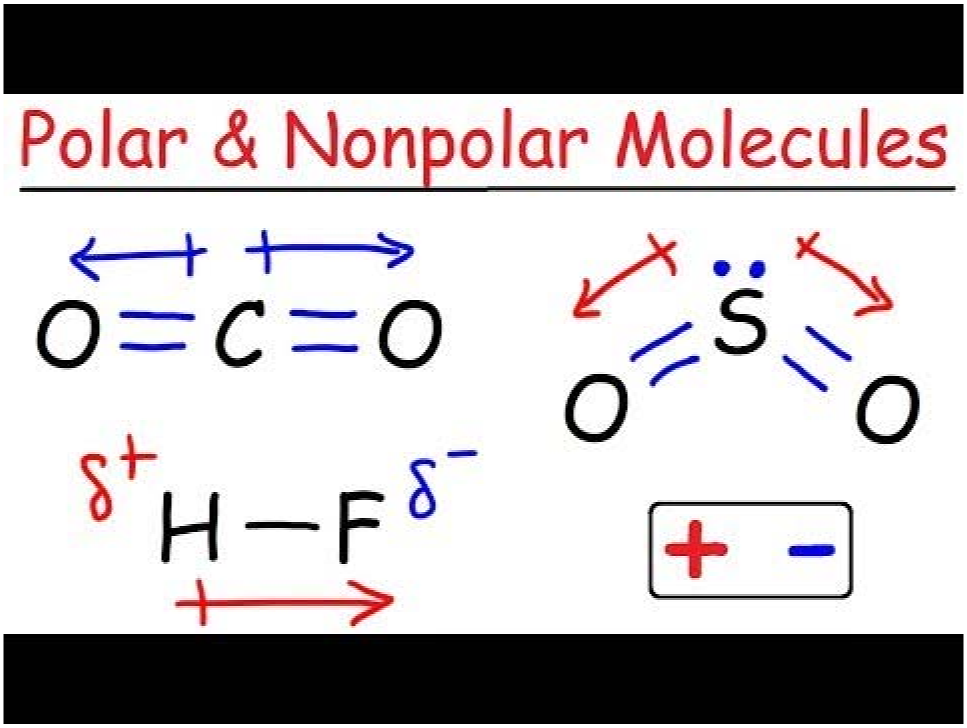
Ans: a) tetrahedral, trigonal pyramidal, less than 109.5
b) Trigonal planar, trigonal planar, 120
c) Tetrahedral, Tetrahedral, 109.5
d) Trigonal planar, angular, less than 120
e) Trigonal planar, trigonal planar, 120
f) Trigonal planar, trigonal planar, 120

XV. Molecular Polarity

Polarity is very important to determine the outcome of most chemical reactions.
In this section polarity of entire molecule will be discussed. A polar molecule is always attracted to the opposite sides of an electric field. (shown in the figure below). Molecular polarity depends on the individual bond polarity as well as symmetry of the molecule.
A polar molecule has either on polar bond or two or more bond dipoles that do not cancel.
A nonpolar molecules has either all nonpolar bonds or two or more polar bonds that do cancel each other.


Electric field
δ+
δ−
−
+
(a)
(b)


O
H
H
C
Polar molecule? yes
Polar molecule? no
(a)
(b)



Complete this Phet simulation Activity below:

https://phet.colorado.edu/en/simulation/molecule-polarity

(This simulation is explained in detail in chapter 7) More information about molecular polairy is given in chapter 7.

The following worksheet has been taken from AACT

Modeling Molecular Polarity

Complete the chart for your given substance.

Substance	**Bond Polarity Difference in Electronegativity**	**Shape**	**Molecular Polarity**
Water(H_2O)			
Formaldehyde (CH_2O)			

Nitrogen (N_2)			
Carbon tetrachloride (CCl_4)			
Ammonia (NH_3)			

Questions

1. Determine the following molecules as polar or nonpolar

a) CH_2Cl2

b) NH_3

c) HF

d) CO_2

Ans: a) polar

b) Polar

c) Polar

d) Nonpolar

<u>Following activity has been taken from AACT</u>

Making Connections between Electronegativity, Molecular Shape, and Polarity

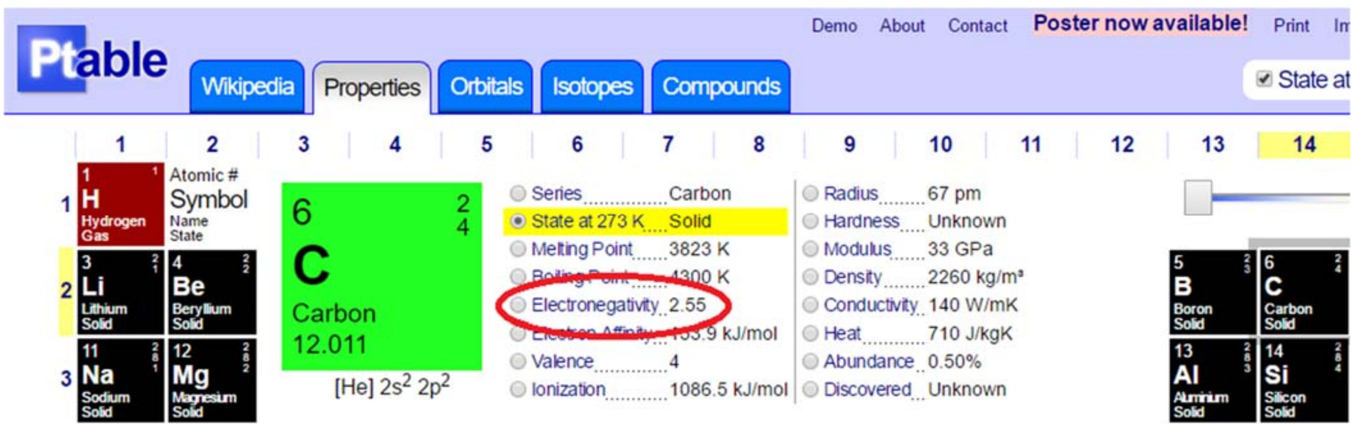
Objective

At the end of this activity you should be able to determine the polarity of a molecule based on the electronegativity of its constituents and its molecular shape.

Directions

Complete each of the following tasks using Ptable.com. Answer the questions as you go along.

1. Go to http://ptable.com/ in your web browser.
2. Select the “Properties” tab at the top.
3. Scroll over an element to see its electronegativity (see image below). Fill in the table below with the electronegativity values for the atoms provided. An example is provided for you.



Element	Electronegativity
C	2.55
H	
Cl	
Br	
N	
O	
F	
P	

4. Which atom is the most electronegative? How can you tell?

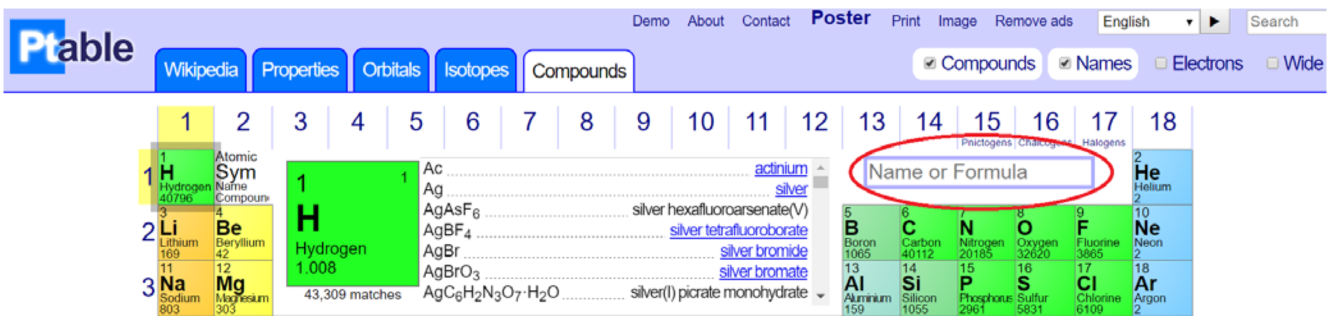
5. If an atom with a high electronegativity value bonds with an atom with a low electronegativity value, what impact does the difference in electronegativity value have on the bond?

6. Draw the Lewis structures of the molecules below and determine their molecular shape:

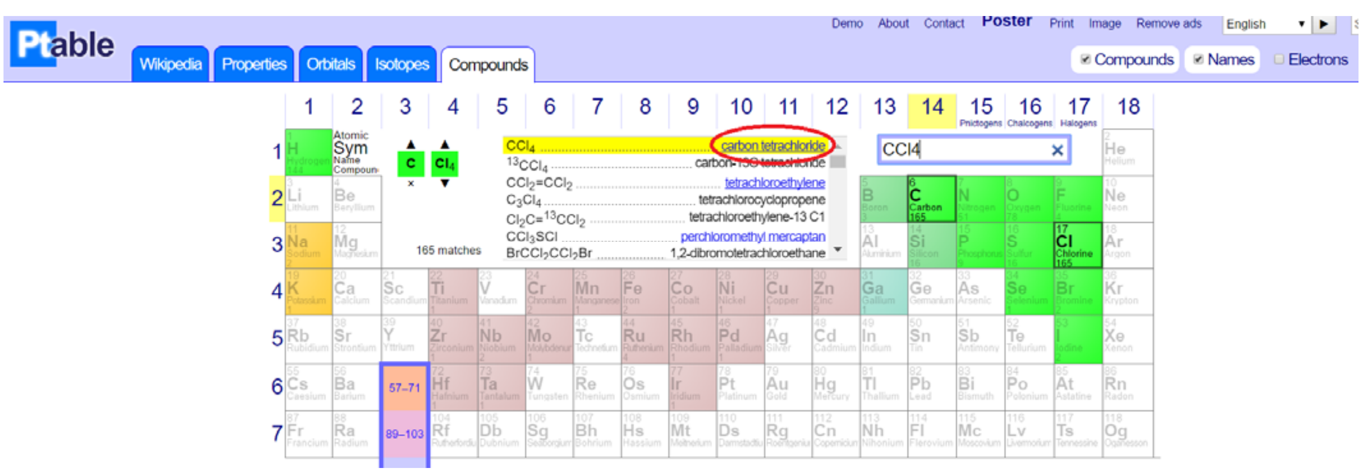
CH_4 Shape:	NH_3 Shape:
CCl_4 Shape:	PCl_5 Shape:
HBr	N_2

Shape:	Shape:
F_2 Shape:	OF_2 Shape:

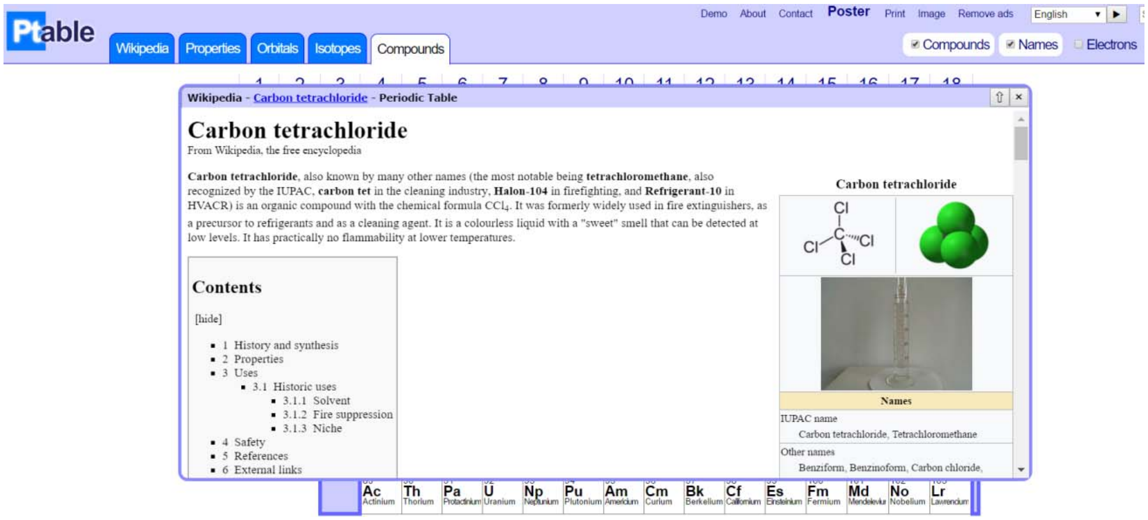
7. Go to the "Compounds" tab on Ptable.com and find the "Name or Formula" search bar. It is circled in the image below.



Search for each of the molecules in question 6. Click the link to the molecule that matches the formula exactly. An example of CCl_4 is shown below.



After you have clicked on the link, a window will pop up on the screen with information regarding the structure and function of the molecule. An example of what you should see for of CCl_4 is shown below. Please note that some structures do not include lone electron pairs.



a. Look at the structure for each molecule and determine if your Lewis structure and shapes were correct. Describe any structures or shapes you had incorrect and explain why they were incorrect:

b. For each molecule, use the <u>*correct shape*</u> in combination with the electronegativity values you found in question 3 to determine if the molecule is polar or nonpolar. Describe how you know if each molecule is polar or nonpolar.

CH_4	NH_3
Polar or Nonpolar? How do you know?	Polar or Nonpolar? How do you know?

CCl_4 Polar or Nonpolar? How do you know?	PCl_5 Polar or Nonpolar? How do you know?
HBr Polar or Nonpolar? How do you know?	N_2 Polar or Nonpolar? How do you know?
F_2 Polar or Nonpolar? How do you know?	OF_2 Polar or Nonpolar? How do you know?

c. Read about each molecule and write down one interesting fact about each molecule.

CH_4	NH_3
CCl_4	PCl_5

HBr	N_2
F_2	OF_2

8. Considering differences in electronegativity values between bonded atoms and the overall molecular shape, what do nonpolar molecules have in common?

9. Considering differences in electronegativity values between bonded atoms and the overall molecular shape, what do polar molecules have in common?

10. Write a short paragraph describing how electronegativity and molecular shape determine the polarity of a molecule. In your paragraph, describe why chemists consider relative electronegativity values and not absolute electronegativity values when determining the

polarity of a bond. Also explain how a molecule with polar bonds can be non-polar overall.

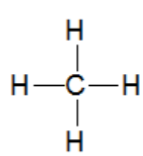
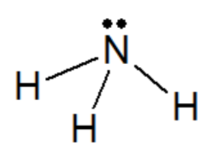
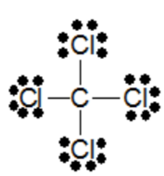
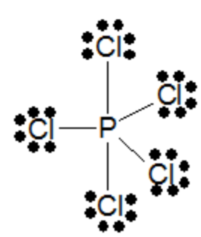
Answer Key: Making Connections between Electronegativity, Molecular Shape, and Polarity

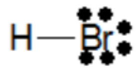
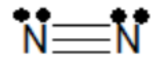
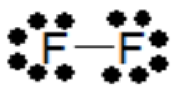
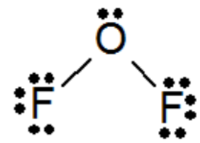
Question 3:

Element	Electronegativity
C	2.55
H	2.20
Cl	3.16
Br	2.96
N	3.04
O	3.44
F	3.98
P	2.19

Question 4:
Fluorine is the most electronegative. You can tell because it has the highest electronegativity number.

Question 5:

CH_4  Shape: Tetrahedral	NH_3  Shape: Trigonal pyramidal
CCl_4  Shape: Tetrahedral	PCl_5 

	Shape: Trigonal bipyramidal
HBr  Shape: Linear	N_2  Shape: Linear
F_2  Shape: Linear	OF_2  Shape: Bent

Question 7b:

CH_4 Polar or Nonpolar? Nonpolar How do you know? All of the bonds in the molecule are nonpolar and the molecule is completely symmetrical.	NH_3 Polar or Nonpolar? Polar How do you know? The bonds in the molecule are polar and the molecule is not completely symmetrical.
CCl_4 Polar or Nonpolar? Nonpolar How do you know? Although all of the bonds in the molecule are polar, the molecule is completely symmetrical, so it is still a nonpolar molecule.	PCl_5 Polar or Nonpolar? Nonpolar How do you know? Although all of the bonds in the molecule are polar, the molecule is completely symmetrical, so it is still a nonpolar molecule.
HBr Polar or Nonpolar? Polar How do you know? The bonds in the molecular are polar and the molecule is not completely symmetrical.	N_2 Polar or Nonpolar Nonpolar How do you know? All of the bonds in the molecule are nonpolar and the molecule is completely symmetrical.

F_2 Polar or Nonpolar? Nonpolar How do you know? All of the bonds in the molecule are nonpolar and the molecule is completely symmetrical.	OF_2 Polar or Nonpolar? Polar How do you know? The bonds in the molecular are polar and the molecule is not completely symmetrical.

Question 7c:
Answers will vary.

Question 8:
Nonpolar molecules can have polar or nonpolar bonds as long as the molecule is symmetrical.

Question 9:
Polar molecules have polar bonds and are non-symmetrical.

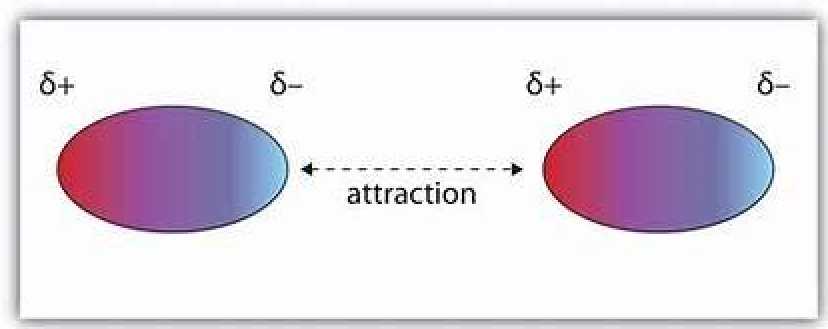
XVI. Intermolecular Forces

The force of attraction that exists between molecules including polar and nonpolar molecules. These forces are responsible for solid, liquid and gas formation. Stronger the IMFs, more tightly the atoms are connected together and more condensed is the phase of the molecule. These forces are so strong that they can lead to phase changes like condensation or solidification.

1. Dipole-dipole force: Exists between any two polar molecules. Opposite poles of molecules attract each other. The attraction between positive end of one dipole and negative end of another dipole is called dipole dipole attraction. For example: two molecules of comparable molecular weight, Butane C_4H_{10} and acetone C_2H_6O, one is polar and another is nonpolar.

$$\overset{\delta^-}{\mathrm{Cl}}-\overset{\delta^+}{\mathrm{H}}\text{ - - - - - - }\overset{\delta^+}{\mathrm{H}}-\overset{\delta^-}{\mathrm{Cl}}$$

2. London dispersion Force: Exists between two nonpolar molecules, formed due to momentary displacement of electrons and protons. Below the example of attraction between two I_2 molecules.



London dispersion forces have their origin in electrostatic interactions. An instantaneous dipole is created within an atom or molecule due to partial separation of electrons from protons. This temporary dipole induces temporary dipole to adjacent molecule and that way can attract another molecule. Nonpolar atom or molecule or molecules usually don’t have any permanent dipole and therefore they are mainly gases with small molar mass but tend to be liquid or gas with increasing molar mass.

When molar mass increases, temporary separation between electrons and protons in atoms and molecules also increases, degree of London dispersion force increases. For example Ne is a monoatomic gas but Xe and Rn can develop attraction with other atoms and form molecular compounds.

Dispersion Forces

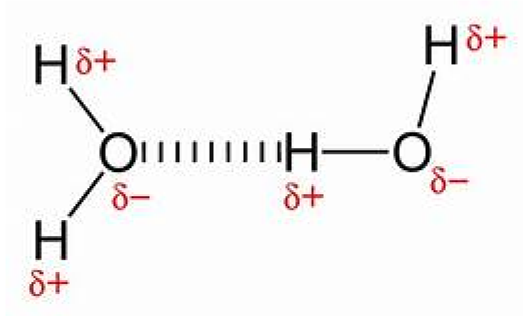
Dispersion forces increase with increasing molecular weight and surface area.

Halogen	molar mass (g)	Boiling Point (K)	Noble gas	molar mass (g)	Boiling point (K)
F_2	38.0	85.1	He	4.0	4.6
Cl_2	71.0	238.6	Ne	20.2	27.3
Br_2	159.8	332.0	Ar	39.9	87.5
I_2	253.8	457.6	Kr	83.8	120.9

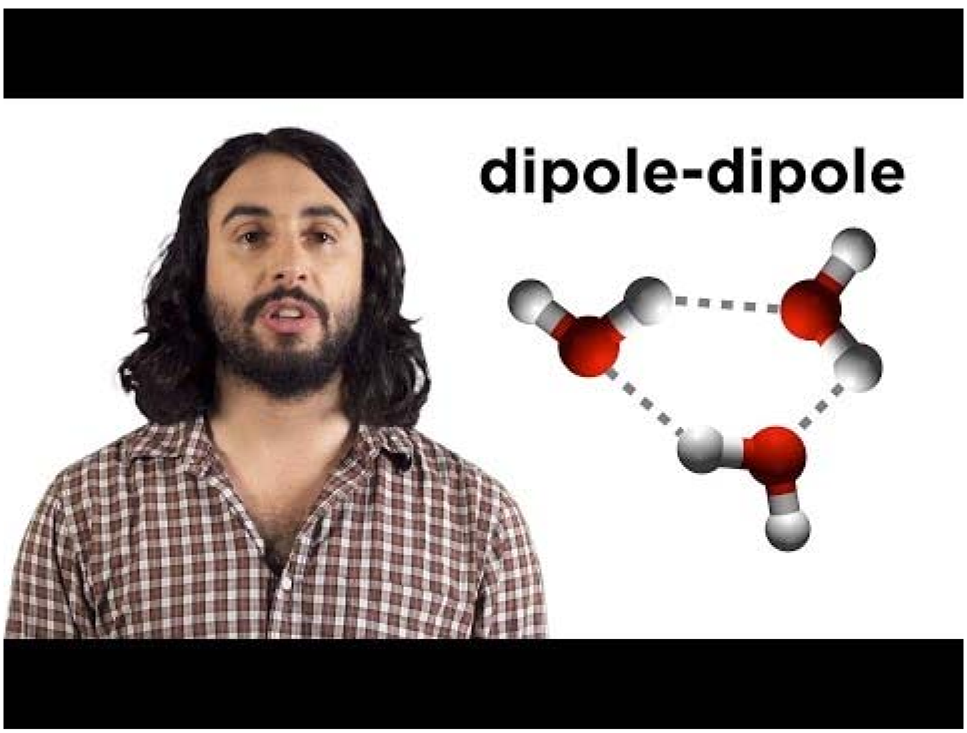
3. H-bonding

When H is present in a polar molecule and attached to a very strong, small electronegative atom like N, F and O, the dipole-dipole force becomes very strong and called H-bonding. This the strongest among all intermolecular forces. The strength of H-bonding ranges from 2 to 10 Kcal/mol. Because of the strong H-bonding, these molecules require extra energy to separate from each other in liquid phase, hence very high boiling point.

Example: H_2O, NH_3, CH_3OH, HF.



https://www.youtube.com/watch?v=08kGgrqaZXA



PRACTICE QUESTIONS: Chapter 4

1. For which of the following atoms do we not try to apply the octet rule?
   a. oxygen b. nickel
   c. sodium d. xenon

2. Which of the following is the stannic ion?
   a. Sn^{+} b. Sn^{2+}
   c. Sn^{3+} d. Sn^{4+}

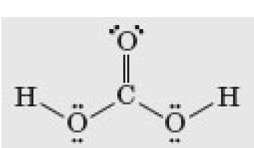
3. A compound used in the treatment of manic depression has the formula Li_2CO_3. Which of the following is the correct name for this compound?
   a. dilithium carbonate b. lithium carbon trioxide
   c. lithium carbonate d. lithium carboxide

4. Which of the following occurs when a sulfur atom is converted to S^{2-}?
   a. The sulfur atom gains two electrons and loses two protons.
   b. The sulfur atom gains two electrons.
   c. The sulfur atom loses two electrons and two protons.
   d. The sulfur loses two electrons.

5. How many electrons are associated with a single bond?
   a. 1 b. 2
   c. 4 d. 6

6. How many resonating structures are possible for ozone (O_3) molecule?
   a. 4
   b. 3
   c. 1
   d. 2

The following represents carbonic acid, an important part of the buffer system in blood.



7. How many total valence electrons were used in constructing this structure?

a. 24

b. 12

c. 17

d. 22

e. none of these

8. Which pair of species is most likely to form an ionic bond?

a. two electrically neutral species

b. two electrically charged species, one positive and one negative

c. two negatively charged species

d. two positively charged species

9. Which of the following is the stannous ion?

a. Sn^{+} b. Sn^{2+}

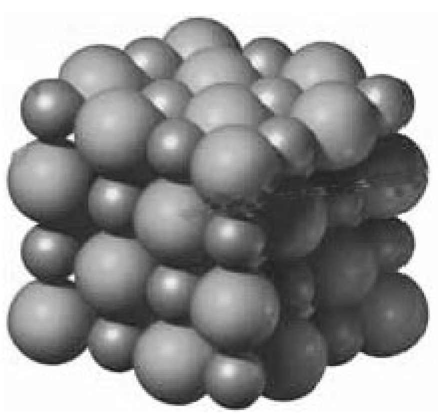
c. Sn^{3+} d. Sn^{4+}

10. What bond angle is most closely associated with a tetrahedral distribution of electron density?

a. 90° b. 109.5°

c. 120° d. 180°

Examine the following image representing the structure of a compound.



11. Which of the following is most likely the type of bonding in this compound?

a. ionic b. nonpolar covalent

c. polar covalent d. electronegative

12. A student is attempting to draw several resonance structures for a species. Which of the following is allowed?

a. Change the number of paired and unpaired electrons.

b. Move electrons from one location to another.

c. both a and b

d. neither a nor b

13. In which of the following molecules does the carbon atom have a linear distribution of electron density?

a. C_2H_2 b. C_2H_4

c. C_2H_6 d. none of these

14. What is the correct name for NH_4NO_3, a compound found in fertilizers?

a. ammonia nitrate b. ammonium nitrate

c. ammonia nitrite d. ammonium nitrite

15. A covalent bond is associated with which of the following?

a. interactions between nuclei b. the sharing of electrons

c. the transfer of electrons d. all of these

16. What is the correct name for KH_2PO_4?

a. potassium hydride phosphate b. potassium hydrogen phosphoxide

c. potassium hydrogen phosphate d. potassium dihydrogen phosphate

17. The ammonium ion has the formula NH_4^+. How many electrons must be shown in the Lewis structure of the ammonium ion?

a. 4 b. 5

c. 8 d. 9

18. In a Lewis structure, which of the following atoms can have more than eight electrons associated with it?

a. C b. O

c. S d. none of these

19. An ionic bond is associated with which of the following?

a. interactions between nuclei b. equal sharing of electrons

c. unequal sharing of electrons d. the transfer of electrons

20. What is the most dominating Intermolecular force present in water?

a. Ion dipole　b. dipole dipole

c. H-bonding　d. dispersion force

21. Which of the following symbols is used to denote that two or more representations of a species are resonance structures?

a. $\rightarrow$　b. $\leftarrow$

c. $\rightleftharpoons$　d. $\leftrightarrow$

22. What is the most stable form of Lithium ion?

a. Li^-　b. Li

c. Li^+　d. Li^{2+}

23. Formaldehyde has the chemical formula CH_2O. How many electrons must be shown in the Lewis structure of formaldehyde?

a. 4　b. 8

c. 12　d. 16

24. What is the name of the species formed when a bromine atom gains an electron?

a. bromate　b. bromide

c. bromine　d. bromine ion

25. Which of the following molecules is(are) polar?

a. H_2O　b. CO_2

c. both of them　d. neither of them

26. Which of the following occurs when a magnesium atom is converted to Mg^{2+}?

a. The magnesium atom gains two electrons and loses two protons.

b. The magnesium atom gains two electrons.

c. The magnesium atom loses two electrons and two protons.

d. The magnesium atom loses two electrons.

27. Which of the following is the correct formula for dinitrogen trioxide?

a. O_2N_3　b. O_3N_2

c. N_2O_3　d. N_3O_2

28. The tin compound used to strengthen teeth has the formula SnF_2. Which of the following is the common name for this compound?

a. stannic fluoride　b. stannous fluoride

c. tin fluoride   d. tin(II) fluorine

29. Which of the following is true of the individual resonance structures which contribute to a resonance hybrid?

a. Each resonance structure actually exists.

b. Each resonance structure is less stable than is the resonance hybrid.

c. In going from one resonance structure to another the electrons move from one location to another.

d. None of the above is true.

30. Which of the following molecules is(are) polar?

a. CH_3Cl b. HCN

c. both of them   d. neither of them

31. What is the formula of the compound formed between the ammonium ion and the carbonate ion?

a. NH_4CO_3 b. $NH_4(CO_3)_2$

c. $(NH_4)_2CO_3$ d. $(NH_4)_3CO_3$

32. What is the common name of Cu^{2+}?

a. copper ion   b. cuprate ion

c. cupric ion   d. cuprous ion

33. Which of the following molecules is(are) polar?

a. $CH_2Cl—CH_2Cl$ b. $CH_3—CHCl_2$

c. both a and b   d. neither a nor b

34. Laughing gas, sometimes called nitrous oxide, has the formula N_2O. What is the proper chemical name for nitrous oxide?

a. nitrogen oxide   b. nitrogen dioxide

c. dinitrogen oxide   d. dinitrogen oxygen

35. In which of the following molecules does the carbon atom have a nearly tetrahedral distribution of electron density?

a. C_2H_2 b. C_2H_4

c. C_2H_6 d. none of these

36. Which of the following is the correct formula for dinitrogen pentoxide?

a. O_2N_5 b. O_5N_2

c. N_2O_5 d. N_5O_2

37. Which of the following elements is most likely to gain electrons to become an anion?

a. Ar b. C

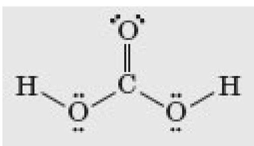
c. Cl d. Fe

38. Which of the following molecules is(are) polar?

a. CH_3Cl b. CH_2Cl_2

c. both of them d. neither of them

The following represents carbonic acid, an important part of the buffer system in blood.



39. How many polar bonds are represented in this structure?

a. 5

b. 3

c. 2

d. None of the bonds are polar.

40. Which of the following atoms is least likely to form an ion?

a. fluorine b. magnesium

c. neon d. sodium

41. The toxic gas hydrogen cyanide has the formula HCN. How many electrons must be shown in the Lewis structure of hydrogen cyanide?

a. 2 b. 6

c. 10 d. 12

42. Formaldehyde has the chemical formula CH_2O. How many nonbonding electrons must be shown in the Lewis structure of formaldehyde?

a. 0 b. 4

c. 8 d. 12

43. Which Greek letter is used in denoting the spatial distribution of charge in a polar bond?

a. alpha, α b. beta, β

c. gamma, γ d. delta, δ

44. The tin compound used to strengthen teeth has the formula SnF_2. Which of the following is the correct systematic name for this compound?

a. stannic fluoride b. stannous fluoride

c. tin fluoride d. tin(II) fluoride

45. Which of the following is generally true of all anions?

a. An anion has equal numbers of protons and electrons.

b. An anion has fewer protons than electrons.

c. An anion has more protons than electrons.

d. An anion has more electrons than neutrons.

46. A student discussing bond polarity and molecular polarity in alcohols made the following statements:

(i) In methanol, CH_3OH, the H—O bond is more polar than the C—O bond.
(ii) Ethanol, CH_3CH_2OH, the alcohol of alcoholic beverages, has at least one polar bond, a net dipole, and is a polar molecule.

Which of the above statements is(are) true?

a. only statement (i) b. only statement (ii)

c. both statement (i) and statement (ii) d. neither statement (i) nor statement (ii)

47. Which of the following is true for the molecule O=C=S?

a. O=C=S is nonpolar because it contains only double bonds.

b. O=C=S is nonpolar because it is linear.

c. O=C=S is polar because it has a bent shape.

d. O=C=S is polar because the polarities of the O=C and C=S do not cancel.

48. Which of the following elements is most likely to lose electrons to become a cation?

a. Ar b. C

c. Cl d. Fe

49. What is the correct name for MgO?

a. monomagnesium monoxide b. magnesium monoxide

c. magnesium oxide d. magnesium oxygen

50. Which of the following is the correct way to write the formula of the compound formed between the barium ion and the sulfate ion?

a. $Ba_2(SO_4)_2$ b. Ba_2SO_4

c. $Ba(SO_4)_2$ d. $BaSO_4$

Answer Key

1. b

2. d

3. c

4. b

5. b

6. d

7. a

8. b

9. b

10. b

11. a

12. b

13. a

14. b

15. b

16. d

17. c

18. c

19. d

20. c

21. d

22. c

23. c

24. b

25. a

26. d

27. c

28. b

29. b

30. c

31. c

32. c

33. b

34. c

35. c

36. c

37. c

38. c

39. a

40. c

41. c

42. b

43. d

44. d

45. b

46. c

47. d

48. d

49. c

50. d