Introduction to Ionic Bonds

The forces that hold matter together are called chemical bonds. There are four major types of bonds. We need to learn in detail about these bonds and how they influence the properties of matter. The four major types of bonds are:

I.	**Ionic Bonds**	**III.**	**Metallic Bonds**
II.	**Covalent Bonds**	**IV.**	**Intermolecular (van der Waals) forces**

Ionic Bonds

The ionic bond is formed by the attraction between oppositely charged ions. Ionic bonds are formed between metals and nonmetals. Remember that metal atoms lose one or more valence electrons in order to achieve a stable electron arrangement. When a metal atom loses electrons it forms a positive ion or **cation**. When nonmetals react they gain one or more electrons to reach a stable electron arrangement. When a nonmetal atom gains one or more electrons it forms a negative ion or **anion**. The metal cations donate electrons to the nonmetal anions so they stick together in an ionic compound. This means that **ionic bonds are formed by the complete transfer of one or more electrons.**

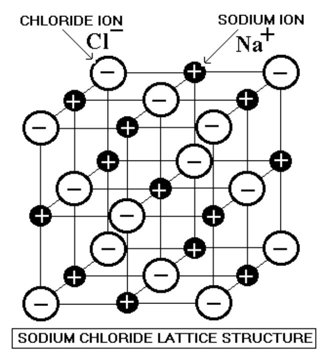
A structure with its particles arranged in a regular repeating pattern is called a **crystal**. Because opposite charges attract and like charges repel, the ions in an ionic compound stack up in a regular repeating pattern called a crystal lattice. The positive ions are pushed away from other positive ions and attracted to negative ions so this produces a regular arrangement of particles where each ion is surrounded by ions of the opposite charge. Each ion in the crystal has a strong electrical attraction to its oppositely charged neighbors so the whole crystal holds together as one giant unit. We have no individual molecules in ionic compounds, just the regular stacking of positive and negative ions.

1. Define the following terms:
   a) ionic bond –

   b) cation –

   c) anion –

   d) crystal –



2. What are the smallest units of an ionic bond?

At room temperature ionic compounds are high melting point solids. They are usually white except for compounds of the transition metals that may be colored. They are brittle (break easily). They do not conduct electricity as solids, but do conduct electricity when melted or dissolved in water.

3. List several properties of ionic compounds:

4. When can electricity to be conducted in an ionic bond?

Reviewing Lewis Dot Diagrams

Write the Lewis Dot Diagrams for the following:

helium atom:

He:

beryllium atom:

Be

beryllium ion:

Be^{+2}

neon atom:

Ne

aluminum atom:

Al

aluminum ion:

Al^{+3}

magnesium atom:

Mg

magnesium ion:

Mg^{+2}

sodium atom:

Na

sodium ion:

Na^{+}

Write the Lewis Dot Diagrams for:

oxygen atom:

O

oxide ion:

$[O]^{-}$

chlorine atom:

Cl

chloride ion:

$[Cl]^{-}$

phosphorus atom:

P

phosphide ion:

$[P]^{-3}$

How would you describe (in general) the Lewis Dot Diagram for:

a) a cation?

Positively charged ion

b) an anion?

Negatively charged ion

What type of bonding would you expect in a compound that contains a metal and a nonmetal?

Ionic bond

Drawing Ionic Bonds

Remember: *Ionic bonds form between POSITIVE IONS and NEGATIVE IONS. Ionic bonding is when one of the atoms is donating an electron(s) (the cation) and one of atoms is accepting an electron(s) (the anion). The electrons are not shared, the anion gains an electron(s) to achieve a full valence and the cation loses an electron(s) to achieve a full valence.*

Diagram the ionic bonding process from neutral atoms to ions showing the valence electrons and indicating with arrows the direction in which the electrons are going. Write your final answer in the box.

Ex: sodium nitride (Na_3N)

Na N

$Na^+[N]^{-3}Na^+$
Na^+

1. sodium chloride (NaCl)

Na Cl

$Na^+[Cl]^-$

5. potassium fluoride (KF)

K F

$K^+[F]^-$

2. barium oxide (BaO)

Ba O

$Ba^{+2}[O]^{-2}$

6. sodium oxide (Na_2O)

Na O

$Na^+[O]^{-2}Na^+$

3. magnesium chloride ($MgCl_2$)

Mg Cl

$[Cl]^- Mg^{+2}[Cl]^-$

7. aluminum chloride ($AlCl_3$)

Al Cl

$[Cl]^- Al^{+3}[Cl]^-$
$[Cl]^-$

4. calcium chloride ($CaCl_2$)

Ca Cl

$[Cl]^- Ca^{+2}[Cl]^-$

8. rubidium oxide (Rb_2O)

$Rb^+[O]^{-2}Rb^+$

Introduction to Covalent Bonds

A covalent bond is formed between nonmetal atoms. The nonmetals are connected by a shared pair of valence electrons. Remember, nonmetals want to gain valence electrons to reach a stable arrangement. If there are no metal atoms around to give them electrons, nonmetal atoms share their valence electrons with other nonmetal atoms. Since the two atoms are using the same electrons they are stuck to each other in a neutral particle called a molecule. **A molecule is a neutral particle of two or more atoms bonded to each other.** Molecules may contain atoms of the same element such as N_2, O_2, and Cl_2 or they may contain atoms of different elements like H_2O, NH_3, or $C_6H_{12}O_6$. Therefore, covalent bonding is found in nonmetallic elements and in nonmetallic compounds.

Covalent bonds are **intramolecular forces**; that is, **they are inside the molecule and hold the atoms together to make the molecule**. Covalent bonds are strong bonds and it is difficult and requires a lot of energy to break a molecule apart into its atoms. However, since molecules are neutral one molecule does not have a strong electrical attraction for another molecule. **The attractions between molecules are called intermolecular forces** and these are weak forces.

Covalent substances have low melting points and boiling points compared to ionic compounds or metals. At room temperature, covalent substances are gases, liquids or low melting point solids. They do not conduct electricity as solids or when molten and usually do not conduct when dissolved in water.

1. Define the following terms:
    a) covalent bond –

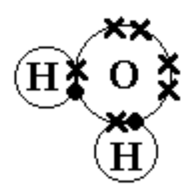
    b) molecule –

    c) intramolecular force–

    d) intermolecular force–

2. List several properties of covalent compounds.

There are many types of covalent bonds. A **single** covalent bond is when two atoms share one pair of valence electrons (see figure). A **double** covalent bond is when two atoms share two pairs of valence electrons. A **triple** covalent bond is when two atoms share three pairs of valence electrons.



3. Define the following terms:
    a) single covalent – 1 pair of e^- being shared

    b) double covalent – 2 pairs of e^- being shared

    c) triple covalent – 3 pairs of e^- being shared

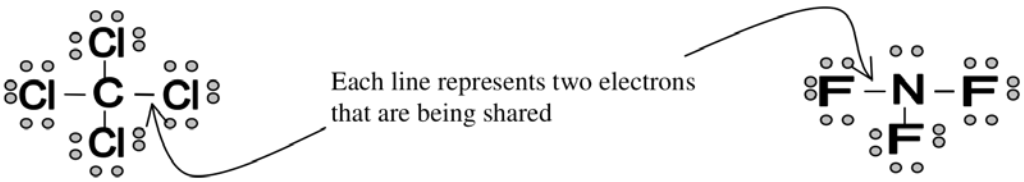
ChemQuest 23

Lewis Structures

Name: ______________________________
Date: ________________
Hour: _____

Information: Drawing Covalent Compounds

For covalent bonding, we often want to draw how the atoms share electrons in the molecule. For example, consider CCl_4 and NF_3 as drawn below:



Notice that the atoms share electrons so that they all have 8 electrons. If you count the electrons around carbon, you will get a total of eight (each line is two electrons). If you count the electrons around each chlorine atom, you will find that there are eight of them.

Critical Thinking Questions

1. How many valence electrons does a carbon atom have (before it bonds)? Hint: find this based on carbon's column on the periodic table.

   4

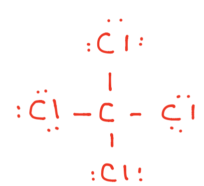
2. How many valence electrons does a chlorine atom have (before it bonds)?

   7

3. Since CCl_4 is made up of one carbon and four chlorine atoms, how many total valence electrons does CCl_4 have? Hint: add your answer to question 1 and four times your answer to question 2.

   32

4. Verify that there are 32 electrons pictured in the drawing of CCl_4.

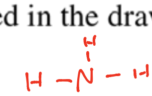
   6 lone pair around each Cl
   4 bonding pairs in the compound

5. Find the sum of all the valence electrons for NF_3. (Add how many valence electrons one nitrogen atom has with the valence electrons for three fluorine atoms.)

   26

6. How many electrons are pictured in the drawing of NF_3 above?

   2

7. In CCl_4 carbon is the "central atom". In NF_3 nitrogen is the "central atom". What is meant by "central atom"?

The atom that is at the center of the compound and surrounded by the other elements and/or lone pairs

8. In SF_3 sulfur is the central atom. You can tell which atom is the central atom simply by looking at the formula. How does the formula give away which atom is the central atom?

least # (only 1 atoms)

9. Identify the central atom in each of the following molecules:

A) CO_2 B) PH_3 C) SiO_2

C P Si

10. For each of the compounds from question 9, add up how many valence electrons should be in the bonding picture. A is done for you.

A) CO_2 B) PH_3 C) SiO_2

4+2(6)=16 $5+3(1)=8$ $4+2(6)=16$

11. The number of electrons that should appear in the bonding picture for CO_3 is 22. The number of electrons that appear in the picture for CO_3^{2-} is 24. Offer an explanation for why CO_3^{2-} has 24 electrons instead of 22. (Where did the extra two electrons come from?)

Taken from another element (metal) to become stable

12. The number of electrons that should be included in the picture of NH_4 is 9. The number of electrons in the picture for NH_4^+ is 8. Offer an explanation for why NH_4^+ has 8 electrons instead of 9.

The "+" means there is 1 more proton than electron. NH_4 has 9 protons, so there has to be 8 e⁻ to have a "+" charge.

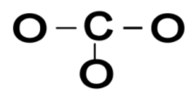
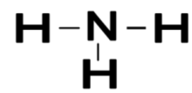
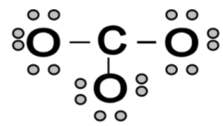
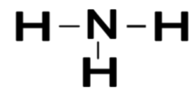
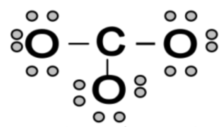
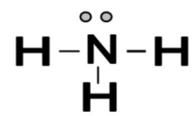
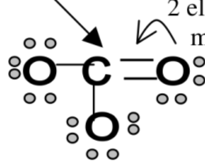
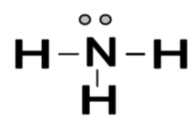
13. Considering questions 11 and 12, we can formulate a rule: For each negative charge on a polyatomic ion, we must add (add or subtract) an electron and for each positive charge we must subtract (add or subtract) an electron.

14. For each of the polyatomic ions or molecules below, determine the total number of valence electrons.

a) NO_3^- 24 b) SCl_4 34 c) H_3O^+ 8 d) PO_4^{3-} 32

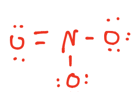
Information: Steps for Drawing Lewis Structures for Covalent Compounds

Study the two examples in the table of how to write structures for CO_3^{2-} and NH_3. Make sure you understand each of the five steps.

	CO_3^{2-}	NH_3
Step #1: Add up the number of valence electrons that should be included in the Lewis Structure.	4 + 3(6) + 2 = **24** (carbon has four and each oxygen has six; add two for the -2 charge)	5 + 3(1) = **8** (nitrogen has five; each hydrogen has one)
Step #2: Draw the "skeleton structure" with the central atoms and the other atoms, each connected with a single bond.		
Step #3: Add six more electron dots to each atom *except* the central atom. Also, never add dots to hydrogen.		 (no change)
Step #4: Any "leftover" electrons are placed on the central atom. Find the number of leftovers by taking the total from Step #1 and subtracting the number of electrons pictured in Step #3.	24 – 24 = 0 leftover electrons  (no change)	8 – 6 = 2 leftover electrons; placed around nitrogen 
Step #5: If the central atom has 8, then you are done. If not, then move two electrons from a different atom to make a multiple bond. Keep making multiple bonds until the central atom has 8 electrons.	a total of 4 electrons are shared here 2 electrons were moved to form a "double bond" 	(no change) 

Critical Thinking Questions

15. Write the Lewis Structure for nitrate, NO_3^{-1}. Hint: when you are done it should look very similar to CO_3^{2-} in the table above.



16. Draw the Lewis Structure for SO_2.

Polarity

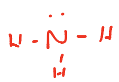
When atoms share valence electrons they do not always share them equally. Frequently one atom has a stronger attraction for the electrons than the other atom does. This uneven attraction causes the electrons to be held closer to one end of the bond than the other; we say this makes one end of the bond slightly positive and the other end of the bond slightly negative. **A covalent bond with uneven sharing of the electrons is called a polar covalent bond. A bond in which the electrons are shared equally is called a nonpolar covalent bond.**

1. Define the following terms:

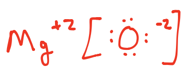
a. polar covalent

b. nonpolar covalent

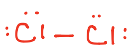
2. Label the following compounds as nonpolar, polar or ionic:

a. NH_3  polar

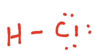
b. MgO ionic



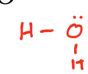
c. Cl_2 non-polar



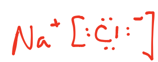
d. HCl polar



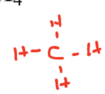
e. H_2O polar



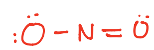
f. NaCl



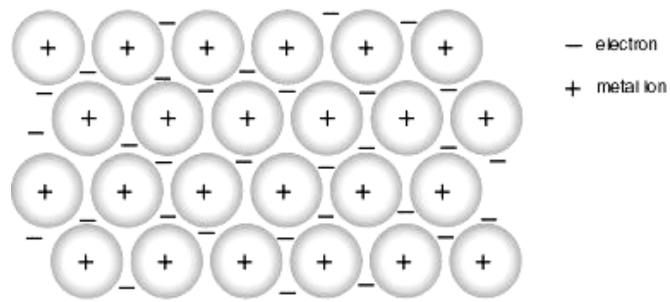
g. CH_4 non-polar



h. NO_2 polar



Metallic Bonds



A metallic bond forms between multiple metal atoms. The metallic bond is formed by the mutual attraction for each others loosely held valence electrons. Most metal atoms have only one or two valence electrons and these are not tightly bound to the atoms. In a piece of metal these valence electrons do not seem to belong to any one of the atoms but are able to move freely through the structure from one atom to another. **Metals can be thought of as positive ions (the nucleus and inner shells of electrons—all of the atom except the valence electrons) in a "sea" of loose valence electrons.** The metal ions line up in a regular repeating pattern (a crystal lattice) and their loose valence electrons move through this crystal acting as an electron glue (see figure). Each of the ions is strongly attracted to all of the loose electrons surrounding it so the whole metal holds together as a crystal.

These electrical attractions for the electron glue are strong and hard to break so metals are high melting point crystalline solids. Since there are charged particles free to move metals are good conductors of heat and electricity as solids and as liquids. Because the "electron glue" is free to move, if we hammer or pull the cations to new positions the electron glue flows right along with the cations and holds the structure together in the new position. Thus, metals are malleable (bendable) and ductile (can be hammered flat) and have a high tensile strength (can be stretched without breaking). This loose cloud of electrons is good at absorbing and re-emitting the light energy that strikes it so metals are lustrous (shiny).

Metallic bonding is found in elemental metals and in **mixtures of metals called alloys.**

1. What is a metallic bond? Explain how the ions and electrons are arranged.

2. List some properties of metallic bonds.

3. What is an alloy?

   Two or more metals bonded together

4. Identify the following compounds as metallic, ionic or covalent:

   a. NaCl Ionic

   b. Cl_2 Covalent

   c. Au Metallic

   d. $[BrO_3]^{-1}$ Covalent (polyatomic ion)

   e. Mg_3N_2 Covalent

   f. Pt Metallic

   g. Al Metallic

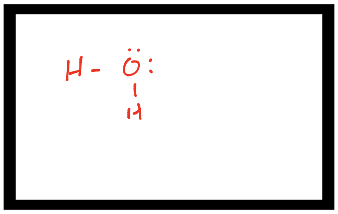
   h. Ag Metallic

Bonding Pictures Review Sheet

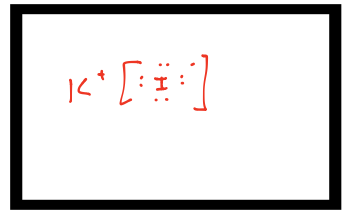
Draw Lewis dot diagrams for the following compounds.

Remember that you must check the difference in electronegativity—to identify the compound as polar or non-polar.

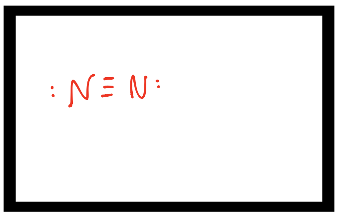
a) water (H_2O)



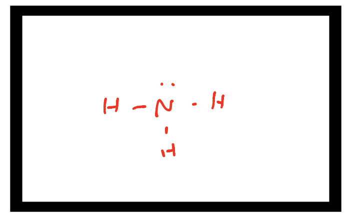
b) potassium iodide (KI)



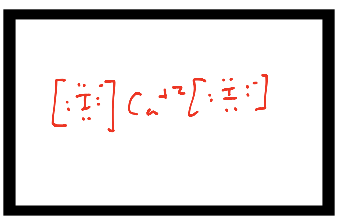
c) nitrogen molecule (N_2)



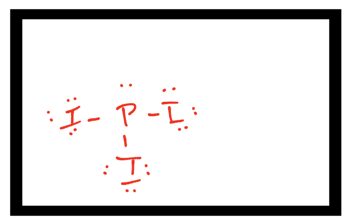
d) ammonia (NH_3)



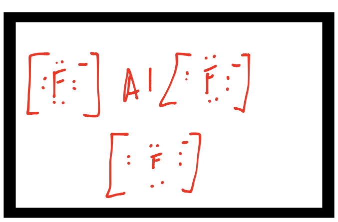
e) calcium iodide (CaI_2)



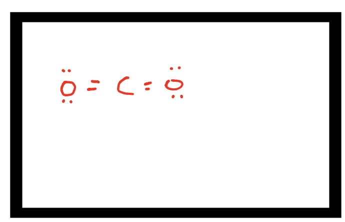
f) phosphorous trichloride (PI_3)



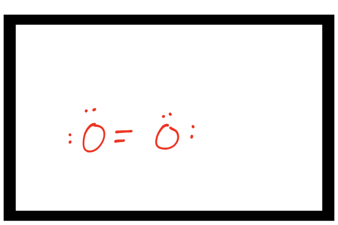
g) aluminum fluoride (AlF_3)



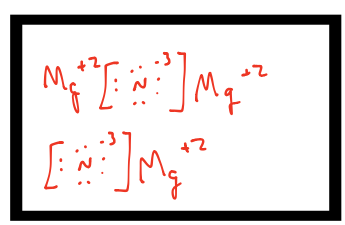
h) carbon dioxide (CO_2)



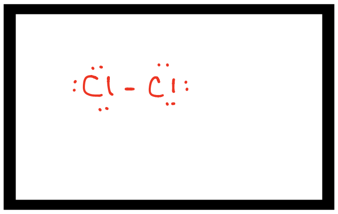
i) oxygen molecule (O_2)



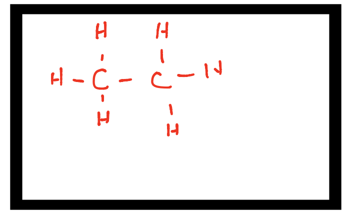
j) magnesium nitride (Mg_3N_2)



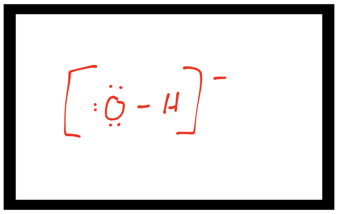
k) chlorine molecule (Cl_2)



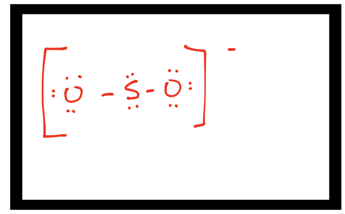
l) ethane (C_2H_6) (carbons hook to each other with H's all around)



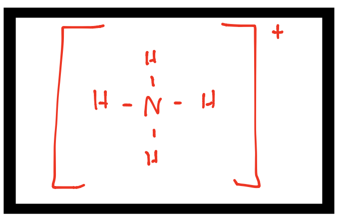
m) hydroxide ion $[OH]^{-1}$



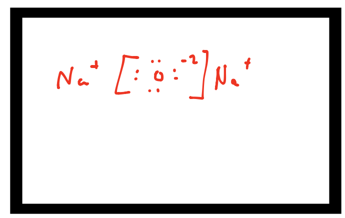
n) sulfite ion $[SO_2]^{-2}$



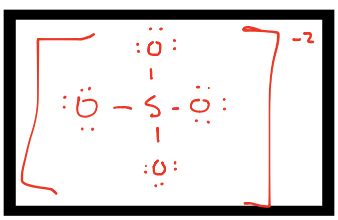
o) ammonium ion $[NH_4]^{+1}$



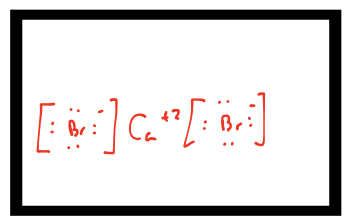
p) sodium oxide (Na_2O)



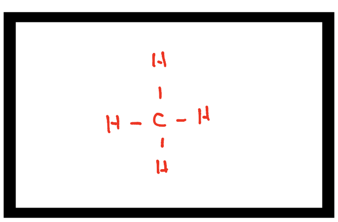
q) sulfate ion $[SO_4]^{-2}$



r) calcium bromide ($CaBr_2$)



s) methane (CH_4)



t) sulfur dioxide (SO_2)