You have mastered this topic when you can:

1) define **IONIC**, **COVALENT** and **POLAR COVALENT BONDS**.
2) use **ELECTRONEGATIVITY** values to predict the type of bond formed between two elements.
3) classify molecular compounds as **COVALENT** or **POLAR COVALENT**.

IONIC AND COVALENT BONDS REVIEWED

I) **IONIC BONDS:** _____

_____.

The transferring of valence electrons from metal atoms to non-metal atoms creates ***metal cations*** (positive ions) and ***non-metal anions*** (negative ions), which are very strongly attracted to one another. The strong attraction between the ***cations*** and ***anions*** is the ***intramolecular force*** called an ***ionic bond***.

II) **COVALENT BONDS:** _____

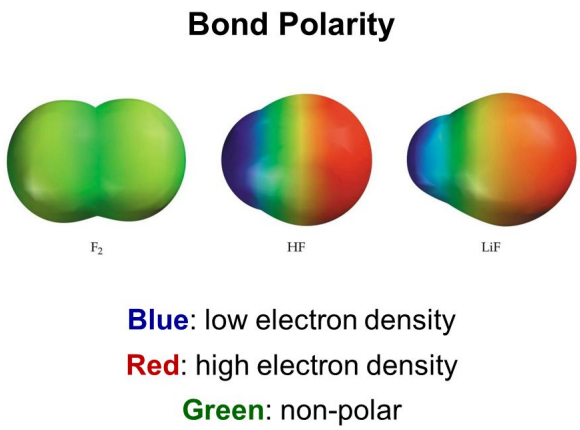
_____.

A) Sharing 2 valence electrons, 1 pair of electrons, creates a single bond.
B) Sharing 4 valence electrons, 2 pairs of electrons, creates a double bond.
C) Sharing 6 valence electrons, 3 pairs of electrons, creates a triple bond.
D) When pairs of valence electrons are shared by atoms, each atom has access to enough valence electrons to give each atom full a valence shell making each atom relatively stable. The attraction the nucleus of each atom has for the shared electrons is the ***intramolecular force*** called a ***covalent bond***.

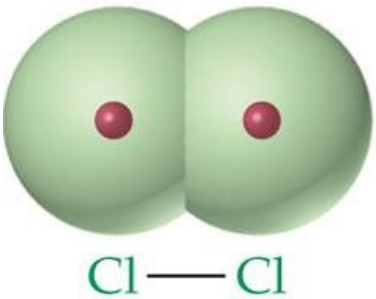
REDEFINING THE COVALENT BONDS

I) When two non-metal atoms share electrons, two possibilities exist: **FIRST:** an ***equal*** sharing of electrons between the two atoms; **SECOND:** an ***unequal*** sharing of electrons between the two atoms.

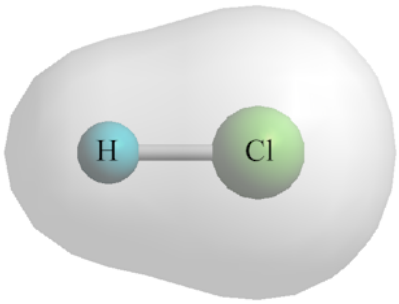
A) Research has revealed that electrons are not distinct particles. Rather, they are forms of matter that exist as both a wave and a particle depending upon the situation they find themselves in. To illustrate this current theory, bonded electrons are represented as grey spheres area around the nucleus of an atom. The size of the sphere around the nucleus of the bonded atoms indicates the relative sharing of the bonded electrons between the atoms. The relative size of the sphere is called the *electron density*. The darker and larger the sphere the greater the *electron density*: the greater the electron density, the greater the likelihood of finding the shared pair of electrons close to the atom. The lighter and smaller the sphere the smaller the *electron density*: the smaller the electron density, the less the likelihood of finding the shared pair of electrons closes to the atom.



1) The diagram to the right shows a $Cl_{2(g)}$ molecule. **NOTICE** that the ***electron density*** around each Cl atom is equal in darkness and size indicating the bond's pair of electrons is shared equally between the two Cl atoms. ***A COVALENT is formed when 1, 2 or 3 pairs of electrons are shared equally between two non-metal atoms.*** Because the electrons are shared equally, the molecule is composed of neutrally charged atoms. ***Covalent bonds are often called NON-POLAR COVALENT.***



2) The diagram to the right shows an $HCl_{(g)}$ molecule. **NOTICE** that the ***electron density*** around the Cl atom is much larger (and is often given a darker colour) than the ***electron density*** around the H atom indicating the bond's shared pair of electrons is closer to the Cl atom than they are to the H atom. This indicates that the bonded pair of electrons is closer to the Cl atom than it is to the H atom. Because the electrons are shared unequally, one end of the molecule, the Cl, has a negative charge while the other end, the H, has a positive charge. A ***covalent bond*** formed as a result of an unequal sharing of a pair of electrons is called a ***POLAR COVALENT BOND*** (often called a ***polar bond***). ***A POLAR COVALENT BOND*** _____



_____.

ELECTRONEGATIVITY AND BONDING

I) **RECALL** that ***electronegativity is a number describing a bonded atom's ability to attract electrons in a bond***. The larger an atom's ***electronegativity***, the greater is its ability to attract electrons in a bond.

A) ***Ionic bonds*** are formed between metal and non-metal atoms. Consider the ***ionic compound*** $NaCl_{(s)}$. Use your periodic table to find the ***electronegativity*** (EN) values of each element: Na: EN = 0.9 while Cl: EN = 3.0. Since Cl has a much larger ***electronegativity***, it has a much greater attraction for electrons, so much so that sodium's single valence electron actually transfers to and becomes part of the chlorine atom creating Na^+ cations and Cl^- anions. The Cl^- ions and Na^+ ions are strongly attracted to each other creating an ***ionic bond***.

B) ***Covalent bonds (non-polar covalent bonds) are formed when two non-metal atoms share electrons equally***. To help understand what a covalent bond is, imagine a pair of electrons attached to a rope that is being pulled by each nucleus of two atoms, each located at opposite ends of the rope.

1) Consider the molecule $Br_{2(l)}$. Use your periodic table to find the ***electronegativity*** (**EN**) value of Br.

Br: EN = 2.8

a) Since each Br atom has an equal ***electronegativity***, each has the same ***attraction*** (pull on the rope) for the shared electrons in the bond, thus the electrons are shared equally by the two Br atoms making the bond ***non-polar covalent***.

2) Consider the molecular compound $CI_{4(l)}$. Use your Periodic Table to find the ***electronegativity*** values of C and I.

C: EN = 2.5 while I: EN = 2.5

a) Since both C and I atoms have equal ***electronegativities***, each has the same ***attraction*** (pull on the rope) for the shared electrons in the bond, thus the electrons are shared equally by the C and I atoms making the bond ***non-polar covalent***.

C) ***Polar Covalent Bonds are formed when two non-metal atoms share electrons unequally***. Consider the molecular compound $HCl_{(g)}$. Use your periodic table to find the ***electronegativity*** values of each element.

H: EN = 2.1 while Cl: EN = 3.0

1) Both H and Cl atoms have relatively large ***electronegativity*** values so they both have fairly strong ***attraction*** (pull on the rope) for the electrons within their bond. Since the chlorine's ***electronegativity*** is larger than that of hydrogen, it has a larger ***attraction***, pull on the rope, for the bond's electrons, thus they are closer to the Cl atom than they are to the H atom. Since the electrons in the bond are closer to, but have not been transferred to the Cl atom, we say the Cl is '*slightly negative*' while the H atom is '*slightly positive*'. The lowercase Greek letter 'δ', which means a '*slight difference*', is used to show the '*slight difference*' in charge between the H and Cl atoms.

D) **Required Practice 1:** Draw the Lewis structures for each molecule and indicate which atoms are slightly positive and slightly negative. **{Answers are on page 4 of these notes.}**

1) $NH_{3(g)}$ 2) $OCl_{2\,(g)}$ 3) $O_{2(g)}$ 4) $HF_{(g)}$ 5) $SiO_{2(g)}$

III) **PREDICTING THE TYPE OF BOND FORMED BETWEEN ELEMENTS USING ELECTRONEGATIVITY**

A) Chemists use the ***difference in electronegativity*** to predict the type of bond formed between two atoms. The ***Difference in Electronegativity (DEN)*** is calculated using this formula.

DEN = Larger EN − Smaller EN

1) By convention:
   a) A ***DEN*** greater than 1.7 means the bond formed is **likely** ionic.

      DEN > 1.7 means an ionic bond will **likely** form.

   b) A ***DEN*** less than 0.3 means the bond formed is **likely** non-polar covalent.

      DEN < 0.3 means a non-polar covalent predicted bond will **likely** form.

   c) A ***DEN*** of and between 0.3 and 1.7 means the bond formed is **likely** polar covalent.

      0.3 ≤ ***DEN*** ≤ 1.7 means a polar covalent bond will **likely** form.

2) **Sample Problems 2**

   1. Predict the type of bond formed between Zn and O.

   The EN values: Zn: EN = 1.6; O: EN = 3.5 → ***DEN*** =

   Since the ***DEN*** is greater than 1.7, the bond formed will likely be ionic.

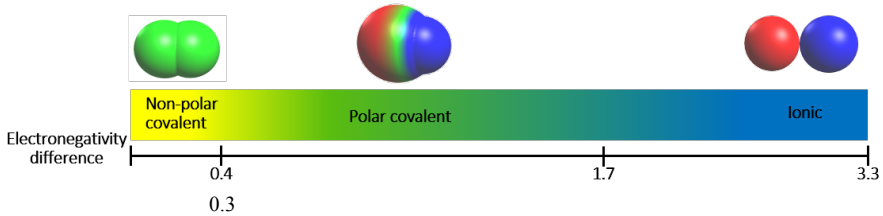
   2. Predict the type of bond formed between N and H.

   The EN values: N: EN = 3.0; H: EN = 2.1 → ***DEN*** =

   Since the ***DEN*** is between 0.3 and 1.7, the bond formed will likely be polar covalent.

B) This model of predicting the type of bond formed between elements has exceptions.

1) **e.g.** $MgI_{2(s)}$ has a ***DEN*** = 1.3, as a result one might predict that it would be held together by ***polar covalent bonds***. Experimental evidence, and the presence of a metal, indicates that its bonds are ionic. This and other exceptions led chemists to develop the ***Bonding Continuum Model***.



   a) The ***Bonding Continuum Model*** states that there is a range of *possible bonds*, which can form between atoms rather than three distinct types. A scale was created where the ***DEN*** is used to place bonds relative to one another: The greater the ***DEN***, the more ionic the bond: the smaller the ***DEN***, the more covalent the bond.

C) Interpreting ***DEN*** using the ***bonding continuum***.

1) ***Ionic compounds***: By definition ***<u>when a compound contains a metal it is an ionic compound</u>*** and is held together by an ***ionic bond*** regardless of the ***DEN***. The ***DEN*** gives an indication of the relative strength of an ionic bond.
   a) **i.e.** The smaller the ***DEN***, the weaker the ionic bond: the larger the ***DEN*** the stronger the ionic bond.
2) ***Molecular compounds***: By definition ***<u>when a compound contains only non-metals it is molecular</u>*** and is held together by ***covalent bonds***. The ***DEN*** is used to determine whether the ***covalent bond*** is ***non-polar covalent*** or ***polar covalent*** and the degree of polarity shown by the ***polar covalent bond***.
   a) When the ***DEN*** < 0.3, the polarity of the bonds created is too small to measure, thus these molecules are classified as ***non-polar covalent***.
   b) When the ***DEN*** ≥ 0.3, the polarity of the bonds created can be measured, thus these molecules are classified as ***polar covalent***: the greater the ***DEN*** the more ***polar*** the bond, the smaller the ***DEN*** the less ***polar*** the bond.

D) ***Electronegativity*** and ***POLAR BONDS***.

1) When the ***DEN*** is between 0.3 and 1.7 ($0.3 \leq$ ***DEN*** ≤ 1.7), the bond formed will likely be ***polar covalent***.

a) Consider the molecular compound $HBr_{(g)}$.

i) Its structural formula is H – Br ; ***Electronegativity values***: H: EN = 2.1; Br: EN = 2.8.

ii) ***DEN*** = _____, thus the bond formed will be _____.
The charges of a ***polar bond*** are indicated by a lower case Greek letter delta δ: **i.e.** $^{\delta+}H - Br^{\delta-}$

b) Consider the molecular compound $NH_{3(g)}$.

i) Its structural formula is H – N – H ; ***Electronegativity values***: H: EN = 2.1; N: EN = 3.0.
|
H

ii) ***DEN*** = _____, thus the bond formed will be _____.

$$^{\delta+}H - N^{\delta-} - H^{\delta+}$$
$$|$$
$$H_{\delta+}$$

c) Consider the molecular compound $CCl_{4(l)}$.

i) Its structural formula is Cl – C – Cl ; ***Electronegativity values***: C: EN = 2.5; Cl: EN = 3.0.
|
Cl
|
Cl

ii) ***DEN*** = _____, thus the bond formed will be _____.

E) **Required Practice 2:** Classify each compound as ionic or molecular, then determine the difference in electronegativity and classify their bonds as ionic, covalent or weakly, moderately or highly polar covalent. **{Answers are on page 4 of these notes.}**

1) $PCl_{3(g)}$ 2) $BrO_{2(g)}$ 3) $SiO_{2(g)}$ 4) $AlN_{(g)}$ 5) $SF_{2(g)}$
6) $CO_{2(g)}$ 7) $CCl_{4(g)}$

ANSWERS TO THE REQUIRED PRACTICE

Required Practice 1 found on page 2

1) $^{\delta+}H - \ddot{N}^{\delta-} - H^{\delta+}$ (with $H^{\delta+}$ below N) **2)** $^{\delta+}$:Cl – O:$^{\delta-}$ (with :Cl:$^{\delta+}$ below O) **3)** $\ddot{O} = \ddot{O}$ **4)** $^{\delta+}H - \ddot{F}:^{\delta-}$ **5)** $^{\delta-}\ddot{O} = Si^{\delta+} = \ddot{O}^{\delta-}$ **6)** $^{\delta+}Al - \ddot{F}:^{\delta-}$

Required Practice 2 found on page 4

1) DEN = 0.9, molecular, the bonds are moderately polar. This is a molecular compound because it is composed of only non-metals and it is moderately polar because the DEN = 0.9. **2)** DEN = 0.7, molecular, the bonds are moderately polar. This is a molecular compound because it is composed of only non-metals and it is moderately polar because the DEN = 0.7. **3)** DEN = 1.7, molecular, the bonds are highly polar. This is a molecular compound because it is composed of only non-metals and it is highly polar because the DEN = 1.7. **4)** DEN = 1.5, ionic, the bonds are ionic. This is an ionic compound because it is composed of a metal bonded to a non-metal. **5)** DEN = 1.5, molecular, the bonds are polar. This is a molecular compound because it is composed of only non-metal atoms. **6)** DEN = 1.0, molecular, the bonds are polar. This is a polar compound because it is composed of only non-metal atoms. **7)** DEN = 0.5, molecular, the bonds are weakly polar. This is a polar compound because it is composed of only non-metal atoms.
