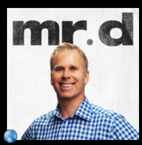
Unit 1: The Diversity of Matter and Chemical Bonding



Nov 17-2:45 PM

Chapter 1: Chemical Bonding

Lesson 1 - Forming Ionic Compounds

Lesson 2 - Forming Covalent Compounds

Lesson 3 - The Nature of Chemical Bonds

Nov 17-2:45 PM

Lesson 1 - Forming Ionic Compounds

In this lesson we will be learning:

Part 1:

1 - How to name ionic compounds.

4 - How to use the periodic table and electron dot diagrams to explain ionic bonding theory.

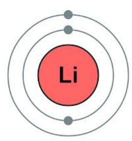
5 - How to explain that an ionic bond results from the simultaneous attraction of oppositely charged ions.

14 - How to use the periodic table to make predictions about bonding and nomenclature.

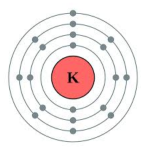
Nov 17-2:45 PM

All elements have a different number electrons in their atoms, that's what makes them unique from one another.

Electrons surround the nucleus of an atom in **shells (energy levels)**. The electrons fill these shells in a particular way. The first (inner most) energy level can only hold **two** electrons, and they will pair together. The second, third, fourth, and so on, can hold a maximum of **eight** electrons each. An energy level must be filled before a new one can be created.



Lithium - 3 electrons



Potassium - 19 electrons

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You can determine the valence (how many electrons in the outer shell) of main group elements by the position of their column in the periodic table.

	Energy Level	Group 1	2	13	14	15	16	17	18
Period 1	1st	1 H							2 He
Period 2	2nd 1st	1 2 Li	2 2 Be	3 2 B	4 2 C	5 2 N	6 2 O	7 2 F	8 2 Ne
Period 3	3rd 2nd 1st	1 8 2 Na	2 8 2 Mg	3 8 2 Al	4 8 2 Si	5 8 2 P	6 8 2 S	7 8 2 Cl	8 8 2 Ar
Period 4	4th 3rd 2nd 1st	1 8 8 2 K	2 8 8 2 Ca						

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A filled outer energy level makes an atom stable, or unreactive.

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He :Ne: :Ar: :Kr: :Xe: :Rn:

Octet Rule: When chemical bonds form, the atoms gain, lose, or share electrons in such a way that they create a filled outer energy level.

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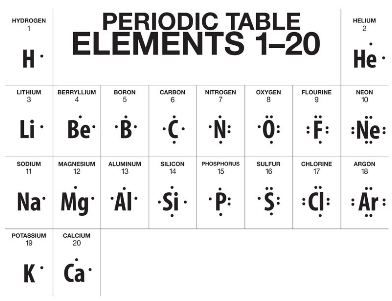
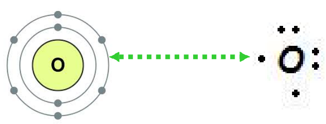
Electron Dot Diagrams are symbolic representations of an atom and its valence electrons. The symbol for the element is used, and dots are placed around it, beginning at the top and going clockwise, spacing out the first four dots equal distance apart. For elements with more than four valence electrons, you place the fifth dot beside the first, and continue clockwise until all of your valence electrons are used.



Example: Oxygen 8 electrons 8-2 = 6 valence electrons



Nov 17-3:05 PM

Forming Ionic Bonds

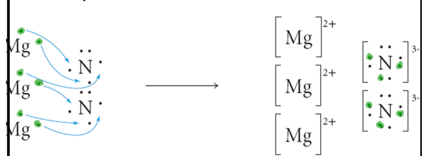
Bond formed by **a metal and a non-metal** through the **transfer** of electron(s).

Metals lose electrons to form **positively** charged ions. **Non-metals** gain electrons to form **negatively** charged ions. Oppositely charged ions attract each other and form ionic bonds.

Writing formulas: Bring the charge on each atom down to be the **subscript on the opposite atom.**

Ex. Mg^{2+} + N^{3-} *Why does nitrogen have a 3- charge?

Mg_3N_2



Nov 17-3:13 PM

Practice! Grab a mini whiteboard.

1. Draw the Electron Dot Diagram for Ca^{2+} and Cl^{-}

2. Draw the Electron Dot Diagram for phosphorous and sodium.

Nov 20-7:28 AM

Naming Ionic Compounds

Nomenclature: rules for naming compounds.

1. Name the metal (+) ion first, followed by the non-metal (-) ion.

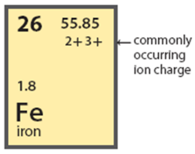
2. The name of the metal ion is the same as the name of the metal atom.

i.e. lithium, sodium, potassium

3. For transition metals that can have more than one possible charge, the magnitude of the charge is indicated Roman numerals in round brackets after the name.

i.e. for Fe^{2+}: iron (IV) ____________

The periodic table gives you the charge of the commonly occurring ions of the transition metals.



4. To name the non-metal ion change the ending of the name of the atom to "-ide".

i.e. chlor**ine** --> chlor**ide**

Nov 17-3:22 PM

Naming Ionic Compounds *with Polyatomic Ions*

Polyatomic Ion - a group of non-metal atoms, covalently bonded together, that carries a charge.

1. Name the **+** ion followed by the **-** ion.

2. The net charge on the compound MUST BE ZERO!

3. If more than one polyatomic ion is needed to make the charge zero, place the formula for the ion inside round brackets, and the **subscript** (for how many you need) outside the brackets. i.e. $(PO_4)_2$

ex. calcium cyanide

Ca^{2+} CN^{-}

$Ca(CN)_2$

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Practice! Grab a mini whiteboard.

Page 13
Questions 3, 4, 7, and 8

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14 - How to use the periodic table to make predictions about bonding and nomenclature.

Nov 17-2:45 PM

Lesson 2 - Forming Covalent Compounds

In this lesson we will be learning:

Part 2:

1 - How to name molecular substances.

2 - Explain why formulas for molecular substances refer to the number of atoms of each element.

3 - How to relate electron pairing to multiple and covalent bonds.

4 - How to draw electron dot diagrams of atoms and molecules, writing structural formulas for molecular substances and using Lewis Structures to predict bonding in simple molecules.

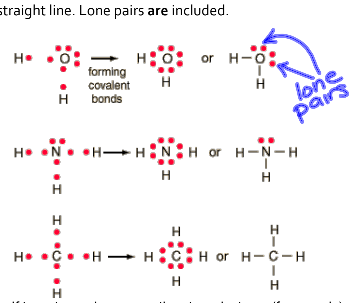
Nov 17-2:45 PM

Forming Covalent (Molecular) Bonds

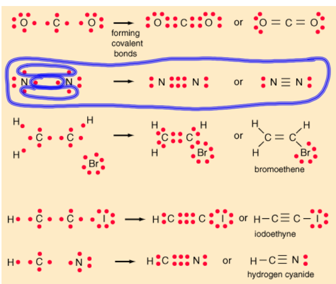
Non-metal atoms **share** electrons to form covalent bonds.

Electron dot diagrams of molecules are called **Lewis structures**.

In structural formulas, a **covalent bond** is represented by a straight line. Lone pairs **are** included.



If two atoms share more than two electrons (four, or six), then represent their structural formula with a double, or triple bond.



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Nomenclature of Covalent Bonds

that is first in the formula first.

Name the atom ~~with the **lowest group number first**~~.

If there is more than one atom, add a **prefix**.

The name remains the same as that of the element.

that is second in the formula, second.

Name the atom ~~with the highest group number second~~.

Add a **prefix**, even if there is only one atom.

Change the ending of the element name to "**-ide**".

Covalent Compounds

The number of each atom is given by prefixes	Mono-	1
	Di-	2
	Tri-	3
	Tetra-	4
	Penta-	5
	Hexa-	6
	Hepta-	7
	Octa-	8
	Nona-	9
	Deca-	10

Ex. CO_2 = carbon **dioxide**

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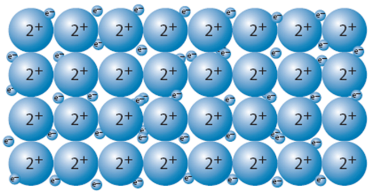
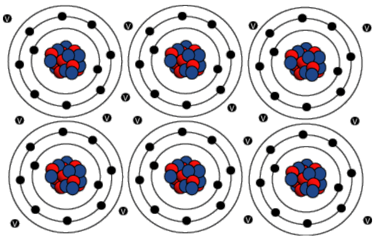
Metallic Bonding

Valence electrons of all atoms are delocalized and can move from one atom to the next. The metal ions are attracted to all nearby electrons.





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Practice!

Grab a mini whiteboard and find a partner :)

1. Use Lewis structures to show the simplest way in which each pair of elements forms a covalent bond, according to the octet rule.
   - (a) hydrogen and oxygen
   - (b) chlorine and oxygen
   - (c) carbon and hydrogen
   - (d) iodine and hydrogen
   - (e) nitrogen and hydrogen
   - (f) hydrogen and rubidium ionic

2. Name each molecule in question 1.

3. One carbon atom is bonded to two sulfur atoms. Use a Lewis structure to represent the bonds.

4. Draw a Lewis Diagram of how bonding occurs between calcium and oxygen.

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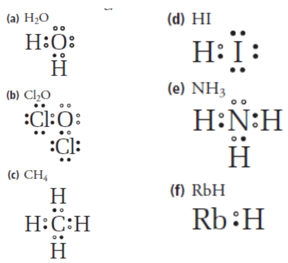
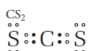
Answers!

1.

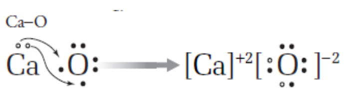


2. a) dihydrogen monoxide
   b) dichlorine monoxide
   c) carbon tetrahydride
   d) hydrogen iodide
   e) nitrogen trihydride
   f) rubidium monohydride

3.



4.



Nov 17-5:49 PM

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Nov 17-2:45 PM

Lesson 3 - The Nature of Chemical Bonds

In this lesson we will be learning:

Part 1:

3 - How to define electronegativity.

Part 2:

10 - How to describe bonding as a continuum ranging from complete electron transfer to equal sharing of electrons.

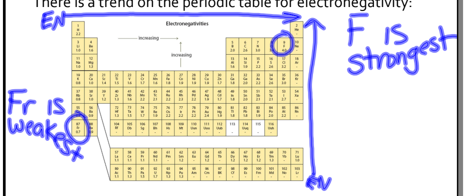
Nov 17-2:45 PM

Electronegativity (*EN*)

> the ability of an element's atoms to attract the shared electrons in a chemical bond

The more electronegative an element is, the **stronger** it is.

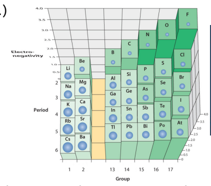
There is a trend on the periodic table for electronegativity:



For main group elements, electronegativity tends to **increase**

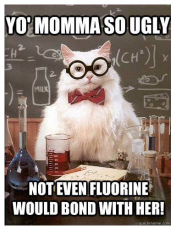
> with the group number (**left to right** on the periodic table)

> as the size of the neutral atom decreases (as you move **up** a group.)



Which is the MOST electronegative element?

Which is the LEAST electronegative element?



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Electronegativity and Bonding

Within each period, as the charge of a nucleus **increases**, the radius of the neutral atom **decreases**.

Non-metals have high electronegativities and metals have low electronegativities, allowing the non-metals to remove electrons from metals and thus form **ionic bonds**.

When their electronegativities are similar in magnitude, atoms of the two elements tend to share electrons in a **non-polar covalent bond**.

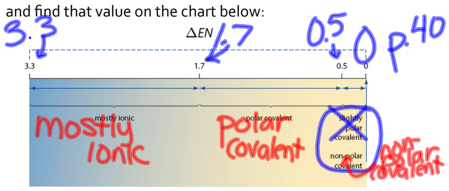
Relatively equal strength

When the electronegativities of two bonded non-metals are quite different, the atom with the higher electronegativity attracts the electrons in the bond more strongly, resulting in a **polar bond**.

Not equal strength

Bonding as a continuum:

To determine the type of bond between two atoms, simply subtract their electronegativities (found on the periodic table), and find that value on the chart below:



We will use 0.5 as our cut-off between non-polar and polar covalent, with 0.5 being polar. When determining between mostly ionic and polar covalent you must look at the TYPE of element as well (metal and non-metal or two non-metals).

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Let's Review!

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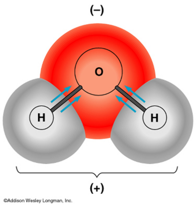
Examples:

1. Water. Which type of bonds are the O-H bonds?

Oxygen - 3.4 Hydrogen - 2.2
Electronegativity difference of 1.2 makes it **polar covalent!**



2. Methane, CH_4. Which types of bonds are C-H bonds?

Carbon - 2.6 Hydrogen - 2.2
Electronegativity difference of 0.4 makes it **non-polar covalent!**

Nov 17-5:19 PM

Practice!
Grab a mini whiteboard and find a partner :)

1. Determine whether each bond below is non-polar covalent, polar-covalent, or ionic.

a) O–H	(b) C–H	(c) Mg–Cl	(d) B–F
e) Cr–O	(f) C–N	(g) Na–I	(h) Na–Br

2. Put **electronegativity, polar-covalent bond**, and **non-polar covalent bond** in your vocabulary book.

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Answers!

1.

(a) O–H, $\Delta EN = EN\text{O} - EN\text{H} = 3.44 - 2.20 = 1.24$, **polar covalent**
(b) C–H, $\Delta EN = EN\text{C} - EN\text{H} = 2.55 - 2.20 = 0.35$, **non-polar covalent**
(c) Mg–Cl, $\Delta EN = EN\text{Cl} - EN\text{Mg} = 3.16 - 1.31 = 1.85$, ionic
(d) B–F, $\Delta EN = EN\text{F} - EN\text{B} = 3.98 - 2.04 = 1.94$, **polar covalent**
(e) Cr–O, $\Delta EN = EN\text{O} - EN\text{Cr} = 3.44 - 1.66 = 1.78$, ionic
(f) C–N, $\Delta EN = EN\text{N} - EN\text{C} = 3.04 - 2.55 = 0.49$, **non-polar covalent**
(g) Na–I, $\Delta EN = EN\text{I} - EN\text{Na} = 2.66 - 0.93 = 1.73$, ionic
(h) Na–Br, $\Delta EN = EN\text{Br} - EN\text{Na} = 2.96 - 0.93 = 2.03$, ionic

Nov 17-5:45 PM

Let's Recap:

Part 1:

3 - How to define electronegativity.

Part 2:

10 - How to describe bonding as a continuum ranging from complete electron transfer to equal sharing of electrons.

Nov 17-2:45 PM

Chapter 1 Quiz Tomorrow!

Good practice to do to prepare:

p. 750
Q# 1, 2, 3, 4, 6, 8, 9, 10, 11,13

Answers are on page 741-742

Nov 17-6:09 PM

Chapter 2: Diversity of Matter

Lesson 1 - 3-D Ionic Structures

Lesson 2 - 3-D Covalent Structures

Lesson 3 - Polarity of Molecules

Lesson 4 - Intermolecular Forces

Lesson 5 - Relating Structures and Properties

Nov 18-5:31 PM

Lesson 1 - 3-D Ionic Structures

In this lesson we will be learning:

Part 1:

2 - Explain why the formulas for ionic compounds refer to the simplest whole-number ratio of ions that result in a net charge of zero.

6 - That ionic compounds form lattices

7 - Identification of everyday processes and products in which ionic compounds are significant, such as the composition of household products and foods in life processes.

12 - I can build models of ionic solids.

Nov 18-5:33 PM

Ionic Structures form Crystals

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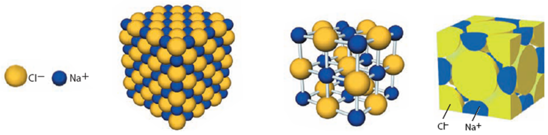
- The packing of ions in a crystal is influenced by the relative sizes and charges of the positively and negatively charged ions.
- Since positively charged ions are attracted to all negatively charged ions, there are no combinations of ions that can be identified as a "molecule."
- This three-dimensional pattern of alternating positive and negative ions is called a **crystal lattice**.



Fun fact!
How many of these are in 1 mg of sodium chloride?

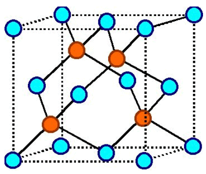
Nov 18-5:36 PM

- Each ion is attracted to **all** of the adjacent ions of the opposite charge.
- The formula $NaCl_{(s)}$ simply means the **ratio** of sodium to chloride ions is one-to-one in the entire crystal.
- CaF_2 means there are two fluoride ions for each calcium ion.
- The smallest ratio of ions in such a crystal is called a **formula unit**, and not a molecule.

Nov 18-5:41 PM

Where do we encounter ionic compounds?

Nov 18-6:03 PM



Activity: Building Ionic Crystals

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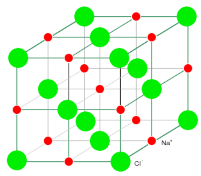
Please complete all steps with a partner, or group of three.
Just discuss the questions for analysis and conclusion, no need to record your answers.



Nov 18-6:00 PM

Let's Recap:

Part 1:

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6 - That ionic compounds form lattices

7 - Identification of everyday processes and products in which ionic compounds are significant, such as the composition of household products and foods in life processes.

12 - I can build models of ionic solids.

Nov 18-5:33 PM

Lesson 2 - 3-D Covalent Structures

In this lesson we will be learning:

Part 2:

5 - How to apply VSEPR theory to predict molecular shapes for linear, bent, tetrahedral, trigonal pyramidal, and trigonal planar molecules.

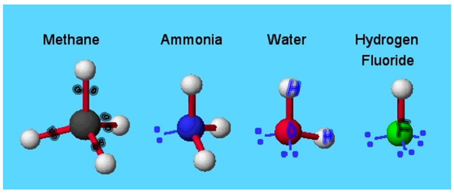
6 - How to illustrate, by drawing or by building models, the structure of simple molecular substances.

15 - How to build models depicting the structure of simple covalent molecules, including selected organic compounds.

Nov 18-6:08 PM

VSEPR = Valence Shell Electron Pair Repulsion Theory.

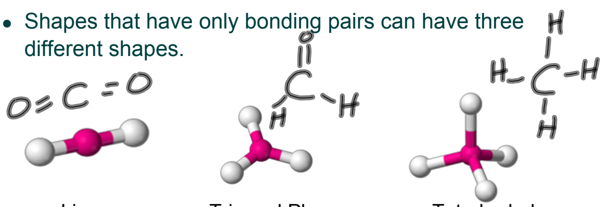
- In 1957, two scientists Ronald Gillespie and Ronald Nyholm developed a model for predicting the shape of molecules.
- The main idea is that **lone pairs** and **bonding pairs** in a molecule **repel each other** because of their negative charges.



- These pairs of electrons will get as far away from each other as they can.
- Lone pairs want to spread out more than bonding pairs.
- So, repulsion is greatest between Lone Pairs (LP-LP).
- Bonding pairs don't repel as much, so the space between bonding pairs (BP-BP) is less than between lone pairs.
- The distance between a BP and a LP will be in between the other two.
- LP-LP > LP-BP > BP-BP
  (greatest) (least)
- Double, triple, or quadruple bonds act as a single electron group when they are repelling other electron groups.

Nov 19-7:30 AM

- By looking at the Lewis structure of a molecule, you can determine its shape.
- Shapes that have only bonding pairs can have three different shapes.

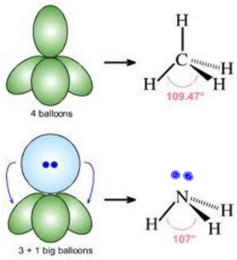


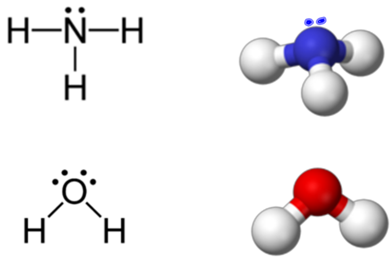
Linear Trigonal Planar Tetrahedral

Let's make models of these!! :)

Nov 19-7:33 AM

- Shapes with lone pairs are more complicated because lone pairs repel more than bonding pairs.
- The angles of the shape are no longer equal.
- Exact angles cannot be predicted.

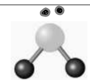
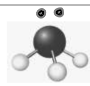
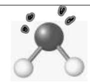




Let's make models of these!! :)

Nov 19-7:37 AM

Molecular Shapes - VSEPR

VSEPR Class	Name of molecular shape	Type of electron pairs	Shape	Example
AX_2	linear	all BP		CO_2
AX_3	trigonal planar	all BP		CH_2O
AX_2E	not popular bent (trigonal planar electron groups)	2 BP, 1 LP		SO_2
AX_4	tetrahedral	all BP		CH_4
AX_3E	trigonal pyramidal (tetrahedral electron groups)	3 BP, 1 LP		NH_3
AX_2E_2	bent (tetrahedral electron groups)	2 BP, 2 LP		H_2O

Nov 20-7:44 AM

VSEPR Link

VSEPR Video

Nov 19-7:40 AM

Steps to Determine VSEPR Shape

1. Draw the Lewis structure
2. Determine the total number of electron groups around the central atom.
3. Determine the types of electron groups (BP or LP).
4. Determine which of the six shapes you have.

Nov 19-7:38 AM

Practice! Grab a mini whiteboard.

1. Briefly describe the primary ideas behind VSEPR theory.

2. For each of the following compounds, a Lewis structure, determine the bond angles and molecular shapes around the central atom:

a) CH_4

b) NH_3

c) CO_2

d) C_2H_2

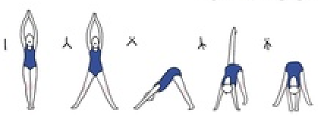
e) CH_2O

Nov 19-7:49 AM

Soap Bubble Activity!

Nov 19-8:01 AM

VSEPR YOGA



Nov 19-7:56 AM

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5 - How to apply VSEPR theory to predict molecular shapes for linear, bent, tetrahedral, trigonal pyramidal, and trigonal planar molecules.

6 - How to illustrate, by drawing or by building models, the structure of simple molecular substances.

15 - How to build models depicting the structure of simple covalent molecules, including selected organic compounds.

Nov 18-6:08 PM

Homework! :)

Page 56
Questions 1-10 (skip number 3)

Nov 19-7:53 AM

Lesson 3 - Polarity of Molecules

In this lesson we will be learning:

Part 2:

9 - I can determine the polarity of a molecule based on simple structural shapes and unequal charge distribution.

Nov 18-6:08 PM

Polar Bonds and Molecular Shapes

If all of the bonds in a molecule are non-polar then the molecule overall is also **non-polar**.

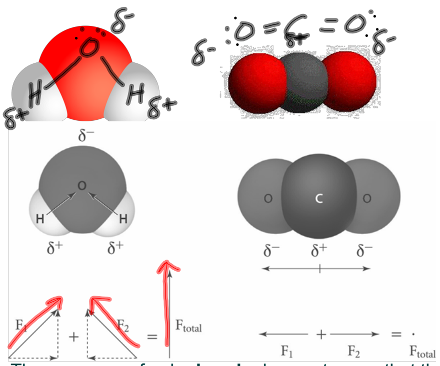
However, if some or all of the bonds are polar, then the molecule **may be** a polar molecule.

In a polar bond, we call the more electronegative atom the "partial negative" atom, represented by **δ -** and the less electronegative atom the "partial positive" atom, represented by **δ +**. This **δ -** and **δ +** is called a **dipole.**

Are all molecules with polar bonds polar molecules?

Let's look at two common molecules to find out.

Water and Carbon Dioxide.



C-2.6
O-3.4
0.8

The presence of polar **bonds** does not mean that the **molecule** is polar.

Nov 21-7:33 AM

Molecular shape	Bond polarity	Molecular polarity
linear	X—A—X	non-polar
linear	X—A—Y	polar
bent	X A X	polar
trigonal planar	X A X X	non-polar
trigonal planar	X A X Y stronger (more EN)	polar
tetrahedral	X X A X X	non-polar
tetrahedral	X X A X Y	polar
pyramidal	X A X X	polar

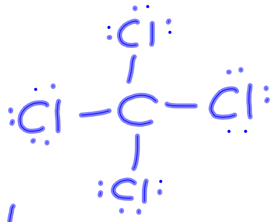
Nov 21-8:30 AM

Practice!

Are each of the following molecules polar?

1. CCl_4
2. $CHCl_3$
3. CCl_2O
4. PCl_3
5. CO_2
6. HCN

hi



ΔEN

C- 2.6

Cl-3.2

0.6 polar bond

Nov 21-8:33 AM

Let's Recap!

Part 2:

9 - I can determine the polarity of a molecule based on simple structural shapes and unequal charge distribution.

Nov 18-6:08 PM

Unit 1 Project - Research a Compound

You may work alone or with one partner on this project. Choose any covalent compound of your choice. Complete internet research and apply concepts learned in class to find the following:

1. Name of your compound - both chemical name and common name
2. Uses of your compound
3. Properties of your compound (3-5 physical properties and 3-5 chemical properties)
4. The electronegativity difference in each bond of your molecule (are the polar or non-polar?)
5. The VSEPR shape of your molecule (or around each 'central' atom in your molecule if it is large)
6. The overall molecule's polarity (polar or non-polar?)
7. What intermolecular forces exist to attract the molecules together
8. 3-5 interesting facts not previously mentioned
9. 3-5 images of your compound (molecular or macroscopic)
10. Build a model of your compound (you may use any material you wish).
11. References (include the website you found information - and pictures! - on) *note: Google Images is NOT a website. Click on the picture in Google Images and find where it comes from.

Nov 22-7:36 AM

Lesson 4 - Intermolecular Forces

In this lesson we will be learning:

Part 2:

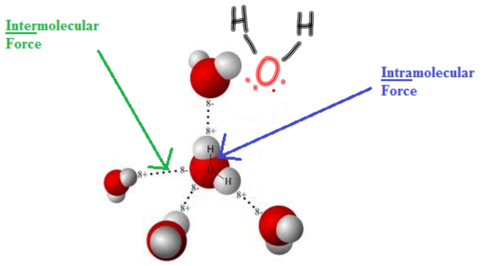
7 - I can explain intermolecular forces, London (dispersion) forces, dipole-dipole forces, and hydrogen bonding.

Nov 18-6:08 PM

Intramolecular forces are the covalent bonds or ionic attractions ***within*** a molecule.

IMF

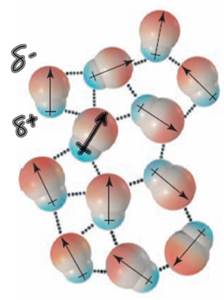
Intermolecular forces are the forces of attraction ***between*** molecules.



Nov 22-7:56 AM

- Polar molecules, possessing dipoles, are attracted to one another by electrostatic attractions between their oppositely charged ends.

p.63

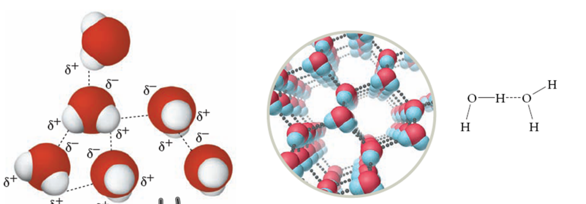


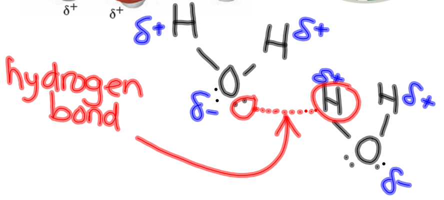
- Dipole-dipole interactions are much weaker than covalent bonds, but they have a significant effect on the structure and function of a compound.

Nov 21-8:45 AM

A very strong dipole-dipole interaction — the hydrogen bond — is critical to the structure of water and many biological molecules.

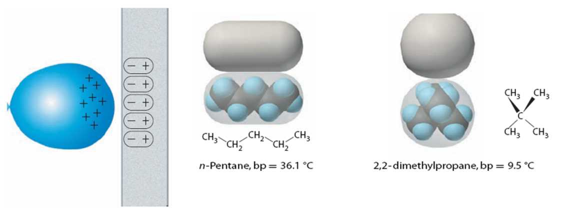
It occurs when a hydrogen atom with a **partial positive** charge is attracted to the **lone pair** on a nitrogen, oxygen, or fluorine atom in a neighboring molecule. Remember NOF.





Nov 21-8:57 AM

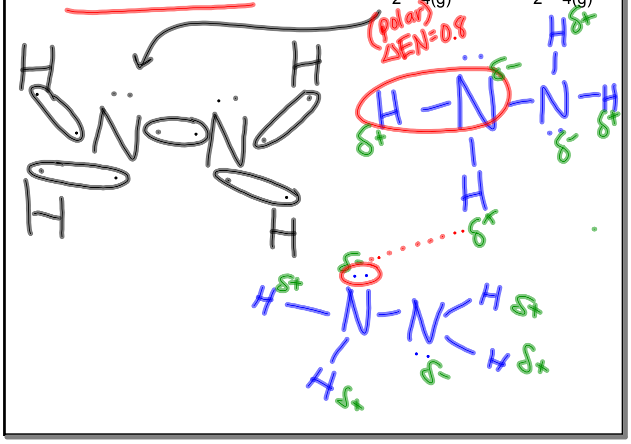
Molecules also interact through London (dispersion) forces, which are attractions between temporary, induced dipoles. In non-polar molecules, these are the only attractive forces.



Nov 21-8:59 AM

Practice! Grab a mini whiteboard :)

1. What is a dipole?
2. Explain how polar molecules interact.
3. What is the difference between dipole-dipole attractive forces and an ionic bond?
4. Discuss the **inter**molecular forces and **intra**molecular forces in $N_2H_{4(g)}$ and $C_2H_{4(g)}$.



Nov 22-7:51 AM

Let's Recap!

In this lesson we will be learning:

Part 2:

7 - I can explain intermolecular forces, London (dispersion) forces, dipole-dipole forces, and hydrogen bonding.

Nov 18-6:08 PM

Lesson 5 - Relating Structures and Properties

In this lesson we will be learning:

Part 1:

6 - How to explain that ionic compounds form lattices and that these structures relate to the compounds properties (i.e. melting point, solubility, reactivity)

8 - How to describe how an understanding of electronegativity contributes to knowledge of relative bond strength, melting points, and boiling points of ionic compounds.

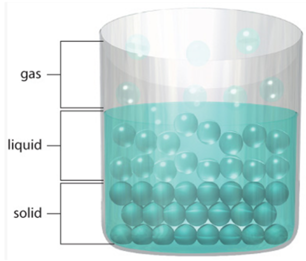
Part 2:

8 - How to relate properties of substances (i.e. melting and boiling points) to the predicted intermolecular bonding in the substance.

11 - How to state a hypothesis and make a prediction about the properties of molecular substances based on attractive forces (i.e. melting point or boiling point).

Nov 18-6:08 PM

The state of matter of a compound is determined by its properties.



Nov 25-8:01 AM

Raising the temperature is raising the kinetic energy (E_k) of particles.

The stronger the appropriate bonding, the higher the mp and bp of the substance.

appropriate bonding:

- **covalents – covalent bonding**
- **ionics – ionic bonding**
- **metals – metallic bonding**
- **intermolecular – LDF, dipole-dipole, H-bonding**

Relative bond strengths:

covalent > ionic > metallic > hydrogen > dipole-dipole > LDF

(LDF varies greatly)

Nov 25-8:07 AM

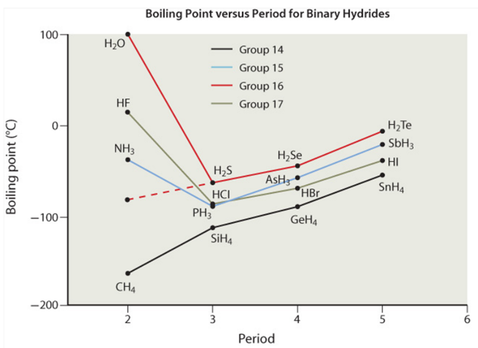
See charts on Page 72.

When ionic or metallic compounds melt or boil, their bonds must be broken.

When covalent substances melt or boil, their intermolecular forces must be overcome (which are much weaker than ionic or metallic bonds). Their actual bonds between atoms need not be broken.

Nov 25-8:09 AM

Hydrogen Bonding has a great effect on boiling point.
(This is on page 74)



Nov 25-8:15 AM

Examples:

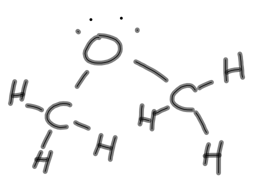
Predict which will have the higher boiling point in each pair:

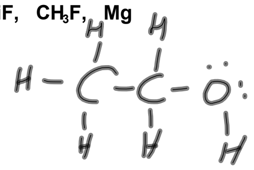
a) C_2H_6 or C_3H_8

b) CH_3F and CH_3OH

Extension: Rank the following substances in order of ↑ bp. State relevant bond types for each

- **CH_3OH, SiC, PH_3, LiF, CH_3F, Mg**





Nov 25-8:20 AM

Note: on your test the substances will always be in pairs

Not everything will necessarily be molecular – i.e. you may get a molecular with an ionic, 2 moleculars; any grouping is possible

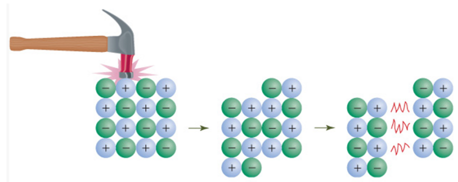
Nov 25-8:22 AM

Why are metals malleable according to the metallic bonding model?
(possible question on next test)

Malleable means workable – bends without breaking; dents without shattering

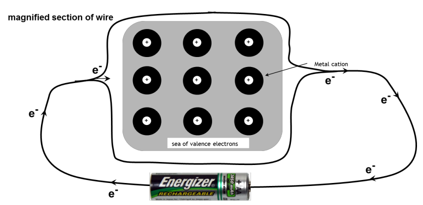
- **metallic bonding model:** ***an array of cations in a sea of valence electrons***
- ***because bonds are between cations and sea of electrons rather than between cations and cations , cations may be moved in the sea of electrons without breaking bonds***

Ionic compounds are not malleable – they shear along **lines as this diagram indicates (page 75)**



Nov 25-8:23 AM

Why do metals conduct electricity according to the metallic bonding model? (possible question on next test)



- **electricity is the flow of electrons**
- **metallic bonding model:** ***an array of cations in a sea of valence electrons***

since valence electrons are loosely held to individual nuclei , they are delocalized – as a new electron enters the wire, another electron must leave to keep wire neutral

Nov 25-8:25 AM

In comparison:
Solid Ionic Conductivity:

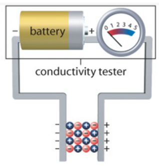


Figure 2.38 Both the positively and negatively charged ions remain in place in a solid ionic compound. No current will register on a conductivity testing meter.

Aqueous Ionic Conductivity:

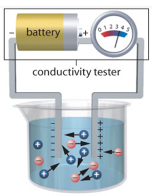


Figure 2.39 An aqueous solution of an ionic compound will conduct electric current because the ions are free to move toward the oppositely charged electrode.

Molecular Compounds Conductivity:
(No - not even if polar)

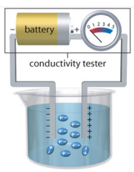


Figure 2.40 If a molecular compound is polar, the positive end will orient toward the negative electrode and the negative end will orient toward the positive electrode. A non-polar molecule will be induced to form a dipole. However, the charges cannot leave the molecule and thus no electric current will flow.

Nov 25-8:27 AM

Practice:
Back of the handout on IMF
Page p. 750 # 15, 17, 18, 20, 21
Answers: P. 741

Lab Tomorrow! :)
Dress appropriately.

Nov 25-8:30 AM

Let's Recap!

Part 1:

6 - How to explain that ionic compounds form lattices and that these structures relate to the compounds properties (i.e. melting point, solubility, reactivity)

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Nov 18-6:08 PM