Parent Guide

Unit 2: Bonding

SC2. Obtain, evaluate, and communicate information about the chemical and physical properties of matter resulting from the ability of atoms to form bonds.

a. Plan and carry out an investigation to gather evidence to compare the physical and chemical properties at the macroscopic scale to infer the strength of intermolecular and intramolecular forces.

b. Construct an argument by applying principles of inter- and intra- molecular forces to identify substances based on chemical and physical properties.

For these standards, the students will design their own lab experiment and then explain the relationship between bonding, forces, and the physical properties they tested in the lab, such as melting and boiling points.

Intramolecular Forces (Bonds)

Bonds are the forces that connect atoms to each other, forming molecules or compounds. There are 3 main types of bonds: ionic, covalent, and metallic. Metallic bonds are the strongest.

Ionic bonds are the next strongest. These compounds form when electrons transfer between elements. This happens because the outermost electron shell of an element prefers to be "full". In the illustration below, hydrogen needs to gain an electron to be full, or lose an electron to be full. Fluorine needs eight electrons to be full so it needs to gain one electron. Because it requires less energy for flourine to gain one electron than to get rid of seven, hydrogen "gives up" its electron to fluorine and the compound becomes neutral. A good way to visualize ions is by drawing the Lewis dot structure for the elements, which shows the valence (outermost) electrons for that atom. Ionic compounds must be neutral.

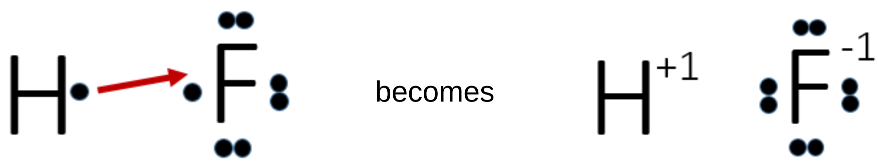


Figure: Lewis structures to show the formation of hydrogen and fluorine ions

Covalent bonds are the weakest, as they are formed by sharing electrons. Electrons can only be shared in pairs between nonmetal atoms. Covalent bonds can either be non-polar (equally shared, weakest type) or polar (unequally shared, the stronger type). Atoms with an electronegativity difference of 0-0.4 are nonpolar, and those with a difference of 0.5 and higher are polar.

Intermolecular Forces

Intermolecular forces attract one molecule to another. The strongest is the ion-dipole force, which occurs when ions are mixed in a polar substance. The next strongest is hydrogen-bonding, a special dipole-dipole force found in polar compounds that contain an H atom AND either an F, O, or N atom. Next is the regular dipole-dipole force, which occurs in all polar molecules. The weakest intermolecular force is the London dispersion force, which is present in ALL substances. The greater the intermolecular forces, the higher a substance's melting and boiling points will be.

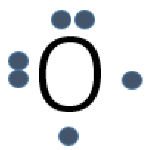
c. Construct an explanation about the importance of molecular-level structure in the functioning of designed materials.

In this standard, students will explore how the structure of a molecule affects the function of man-made materials. Some good research topics would include the chemistry of soap or of metal cookware.

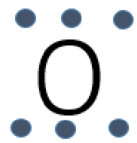
d. Develop and use models to evaluate bonding configurations from nonpolar covalent to ionic bonding.

In this standard, students will both create and use models such as Lewis dot diagrams and Pauling electronegativity charts to determine whether a formula is ionic, covalent, or acidic.

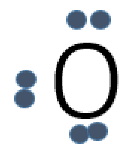
To draw a Lewis dot diagram for an element, first draw the element symbol in the center. Then, draw one dot for each valence electron that element has, following these rules: 1) Dots may only go on the 4 sides [top, bottom, left, right] and 2) Dots cannot be paired until all 4 sides have been used. See below for an example of both correct and incorrect ways to draw the dot structure:



Correct: all 4 sides are used, leaving 2 electron pairs



Incorrect: dots are drawn on the corners



Incorrect: the right side was not used before pairs were made

e. Ask questions about chemical names to identify patterns in IUPAC nomenclature in order to predict chemical names for ionic (binary and ternary), acidic, and inorganic covalent compounds.

In this standard, students will discover the patterns in nomenclature to help them correctly name many types of compounds. In order to successfully name compounds, students must first be able to correctly categorize the formula as ionic, covalent, or acidic. In general, this chart will help your student with these categories:

Type of compound	**The first element is….**
Ionic	A metal or polyatomic ion
Acidic	Hydrogen
Covalent	A nonmetal (excluding hydrogen)

For ionic compounds, the basic patterns/rules for naming are:

1. Write the cation's (positive ion) name first.
2. Change the anion's (negative ion) name ending to –ide.
3. If any ion is polyatomic, instead use the name from the polyatomic ion chart (no changes!).
4. If the cation has a variable charge, then roman numerals must be given in the cation's name to indicate the charge.

For example:

Na_2S = sodium + sulfur = sodium (cation) sulfide (anion with -ide ending) = sodium sulfide

$(NH_4)_2SO_4$ = ammonium + sulfate = ammonium (polyatomic cation) sulfate (polyatomic anion no change) = ammonium sulfate

FeO = iron + oxygen = iron (cation with variable charge) oxide (anion with –ide ending) = iron(II) oxide (charge determined by what is needed to neutralize the anion's -2 charge)

For acidic compounds, the basic patterns/rules for naming are:

1. If binary (containing only H and one other element) use the prefix hydro- and the suffix –ic
2. If oxyacid (containing H and a polyatomic ion) use only the polyatomic ion's name with the new ending –ic (if the polyatomic name ends in –ate) or –ous (if the polyatomic name ends in –ite).

3. End the name with the word "acid"

For example:

$H_2F_{(aq)}$ = hydrogen + fluoride = binary acid = hydrofluoric acid

H_2CO_3 = hydrogen + carbonate ion = oxyacid (change –ate to –ic) = carbonic acid

For covalent compounds, the basic patterns/rules for naming are:

1. Write the first element's name.
2. Change the ending of the second element to –ide.
3. Add a Greek prefix to both the first and second element to tell how many of that atom is present in the compound. *The exception: mono is not written in front of the first element.

The first ten Greek numerical prefixes are shown below:

1	2	3	4	5	6	7	8	9	10
Mono-	di-	Tri-	Tetra-	Penta-	Hexa-	Hepta-	Octa-	Nona-	Deca-

Table : Greek prefixes for 1-10

For example:

CCl_4 = 1 carbon + 4 chlorine = monocarbon + tetrachloride (add prefixes, change to –ide ending) = carbon tetrachloride (no mono for first element)

N_2O_3 = 2 nitrogen + 3 oxygen = dinitrogen trioxide

f. Develop and use bonding models to predict chemical formulas including ionic (binary and ternary), acidic, and inorganic covalent compounds.

For this standard, students will use Lewis dot structures and oxidation states to predict the most likely chemical formula for a compound. For both ionic and acidic compounds, the compound's formula is predicted based on the charges of the two ions, knowing that the total compound must equal zero. The easiest way to do this is to "switch" the charge numbers and reduce to the lowest whole numbers. This is then written as subscripts.

For example:

Rubidium oxide = Rb and O = Rb^{+1} and O^{-2} = Rb_2O_1 = Rb_2O

Titanium(IV) oxide = Ti and O = Ti^{+4} (because of the roman numeral) O^{-2} = Ti_2O_4 (can be reduced) = TiO_2

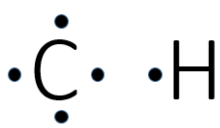
Covalent compound formulas can be found in their names (from the prefixes) or can be predicted based on modeling using Lewis dot structures.

From name prefixes:

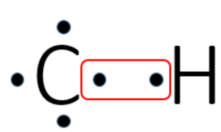
Tetraphosphorus monosilicide = 4 P and 1 Si = P_4Si

From Lewis dot structures:

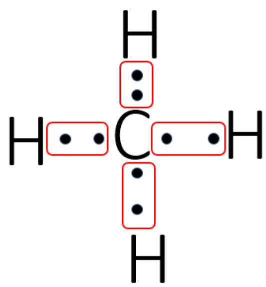
"Carbon bonds with hydrogen. Predict the formula of this covalent molecule."



Draw the lewis dot structures for the individual elements



Create shared pairs



Add more of either atom as needed to make both elements have a full outer shell (C needs 8, H needs 2 electrons).

1 C and 4 Hs = the predicted formula is CH_4

g. Develop a model to illustrate the release or absorption of energy (endothermic or exothermic) from a chemical reaction system depends upon the changes in total bond energy.

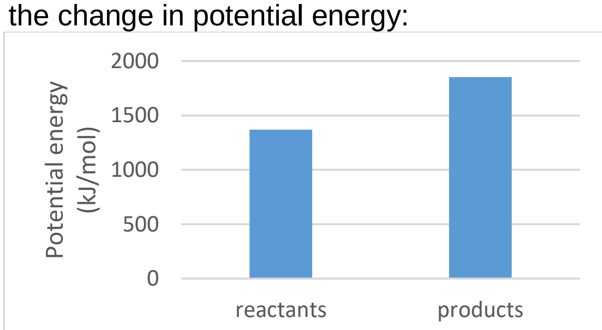
For this standard, students will need to calculate the total bond energy of a reaction based on bonds formed and bonds broken. Then, students can use this to create model of the change in potential energy. If the amount of potential energy increases, then the reaction is endothermic (absorbing energy). If the amount decreases, then the reaction is exothermic (releasing energy).

In order to calculate the total bond energy, the student must look up the bond energies for each bond present in every reactant and every product.

For example:

Given the reaction, $2H_2 + O_2 \rightarrow 2H_2O$, model the changes in bond energy to illustrate the energy of this reaction. Reactants are before the arrow; products are after the arrow.

1. Determine the types of bonds present
   H_2 has an H-H bond, O_2 has an O=O bond, and H_2O has 2 O-H bonds
2. Determine the number of each type of bond present
   There are 2 total H-H bonds, 1 total O=O bond, and 4 total O-H bonds (2 H_2Os = 2x2 O-H bonds)
3. Calculate the bond energies from a table of bond energies
   Reactants: 2 H-H bonds (2x 436 kJ/mol) and 1 O=O bond (1 x 498kJ/mol)
   Products: 4 O-H bonds (4 x 463 kJ/mol)
4. Total the energies of the reactants, and separately total the energy of the products
   Reactants: (2 x 436 kJ/mol) + (1 x 498 kJ/mol) = 1370 kJ/mol
   Products: 4 x 463 kJ/mol = 1852 kJ/mol
5. Create a graph to model the change in potential energy:



This particular example shows an endothermic reaction, as energy increased from reactant to product.