ACTIVITY 8– HOW ATOMS INTERACT WITH EACH OTHER

Background Information

Chemical bonds are the forces that hold together atoms in molecules or ions in crystal compounds. When atoms combine, they combine in order to achieve a lower stable energy state. If we speak in general terms we say that atoms can either share electrons or transfer electron(s) between atoms. If the electrons are shared between the atoms, the bond is said to be a covalent bond and the compound is called a molecular compound. If the electrons are transferred from one atom to another atom the compound is called an ionic compound. It should be emphasized that we do not call an ionic compound an ionic molecule.

Covalent Molecules

When hydrogen atoms combine they share their valence electrons forming a covalent bond and a hydrogen molecule results. This is an example of a molecular compound. When carbon shares its four valence electrons with four hydrogen atoms, CH_4 (methane) forms. The carbon atom in CH_4 now has eight electrons around it, satisfying the octet rule. This stable octet of electrons is the characteristic arrangement of the noble gases (except helium, which can only accommodate two electrons in its single energy level). The **octet rule** states that an atom will either give up or accept an appropriate number of electrons in order to completely fill its outermost energy level.

Other examples of common molecular compounds include carbon dioxide, carbon monoxide, water, NH_3 (ammonia), and hydrogen chloride.

Simple binary covalent compounds (composed of only two elements) are formed between two nonmetals. They are named as follows: the first element in the formula is named first, using the full element name. The second element is named as if it were an anion. Prefixes (mono-, di-, tri-, tetra-, etc.) are used to denote the numbers of each element present. We do not use the prefix mono- for naming the first element.

Simple binary covalent compounds

CO	Carbon monoxide
CO_2	Carbon dioxide
CCl_4	Carbon tetrachloride
P_4O_{10}	Tetraphosphorous decaoxide

Ionic Compounds

Ionic compounds result from a transfer of electrons from one element to another. This electron transfer creates **ions** (an atom or group of atoms that carries an electrical charge), which are held together by an electrostatic (attraction between positively and negatively charged species) force. The electrostatic force is very strong and as a result, ionic compounds have relatively high melting points. Sodium chloride (NaCl) and sodium hydroxide (NaOH) are ionic compounds.

Elements from group IA of the periodic table are metals that form positively charged ions (**cations**) with a charge of 1+. Elemental metals from group IIA form ions with a charge of 2+. These ions form as a result of the metals giving up one and two electrons, respectively, making them isoelectronic with their nearest noble gas.

On the right side of the periodic table in group VIIA we have the halogens (F, Cl, Br, I), which tend to accept an electron to form negatively charged ions (**anions**) with a 1- charge. Sodium, a highly poisonous metal, gives up an electron to chlorine, a highly poisonous gas, which accepts the electron and ordinary table salt forms. The sodium becomes a cation, the chlorine becomes an anion, and they are held together as an ionic compound.

When magnesium forms a compound with chlorine, the formula is $MgCl_2$; Mg forms a dipositive cation, so 2 chloride ions are needed in order to form an overall neutral compound.

Simple ionic compounds are named by first listing the positive ion (as the name of the element; Na^+ is called sodium) followed by the anion (by taking the root of the element name and adding *–ide*; Cl^- is called *chloride*). Subscripts are used to denote multiple ions in the formula.

Sometimes ions exist as a collection of elements. These are known as **polyatomic ions**. When you place an ionic compound that contains a polyatomic ion in an aqueous solution, the polyatomic ion stays together as a unit after dissociation. For example, if you dissolve some KNO_3 (potassium nitrate) in water, the compound dissociates into K^+ and NO_3^- ions.

Some common polyatomic ions

Ion	Name	Examples	
NH_4^+	ammonium	NH_4Cl	Ammonium chloride
CO_3^{2-}	carbonate	Li_2CO_3	Lithium carbonate
HCO_3^-	hydrogen carbonate (bicarbonate)	$NaHCO_3$	Sodium hydrogen carbonate
CN^-	cyanide	KCN	Potassium cyanide
OH^-	hydroxide	$Ca(OH)_2$	Calcium hydroxide
NO_2^-	nitrite	$NaNO_2$	Sodium nitrite
NO_3^-	nitrate	$Fe(NO_3)_2$	Iron (II) nitrate
PO_4^{3-}	phosphate	$Ba_3(PO_4)_2$	Barium phosphate
SO_3^{2-}	sulfite	$BeSO_3$	Beryllium sulfite
SO_4^{2-}	sulfate	K_2SO_4	Potassium sulfate

Iron forms two types of ions Fe^{2+} and Fe^{3+}, so the type of ion must be specified.

Electron Dot Structures

In 1916, G. N. Lewis created a simple model that described chemical bonding of the elements. We use his electron dot diagrams (sometimes called Lewis structures) today to accurately represent many compounds. What follows is a simple set of guidelines for writing these diagrams:

1. Determine the total number of valence electrons in the molecule or ion. If dealing with a negative ion (anion), add one electron for each negative charge on the ion. If dealing with a positive ion (cation), subtract one electron for each positive charge on the ion.

2. Determine the connectivity of the elements in the molecule or ion and connect each of the atoms with single bonds.

3. Distribute the remaining valence electrons (in pairs) in such a way as to satisfy the octet rule for all atoms (exceptions; hydrogen needs only two electrons, boron is satisfied with six).

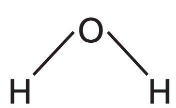
4. Use multiple bonds where necessary: a double bond indicates that the atoms involved share four electrons, a triple bond indicates the sharing of six electrons.

5. Note that a pair of electrons in a covalent bond (a bonding electron pair) can be represented by two dots or a dash. An unshared electron pair (called nonbonding electrons, or a lone pair) can also be represented by two dots or a dash.

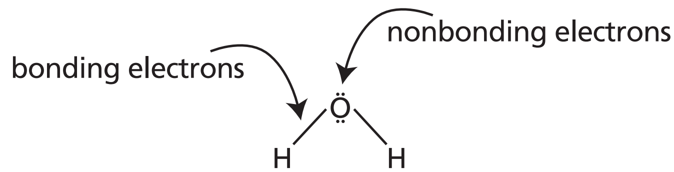
Let's use water as a simple example:

Step 1:
1 oxygen atom × 6 valence electrons = 6
2 hydrogen atoms × 1 valence electron = 2
Total valence electrons: 8

Step 2:
Hydrogen typically forms 1 bond, oxygen typically forms 2 bonds, so to connect them:



Step 3:
We used 4 electrons to connect the atoms; there are four remaining. Hydrogen can take no more than 2 electrons, so we put the last 4 valence electrons (in pairs) on oxygen (as nonbonding electrons). The oxygen atom now has eight electrons around it.



In order to write electron dot diagrams for ions, one must be aware of the concept of formal charge.