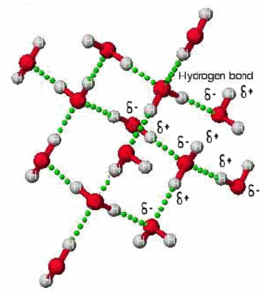
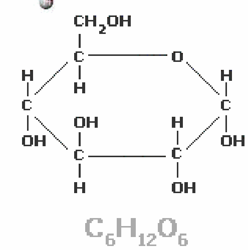
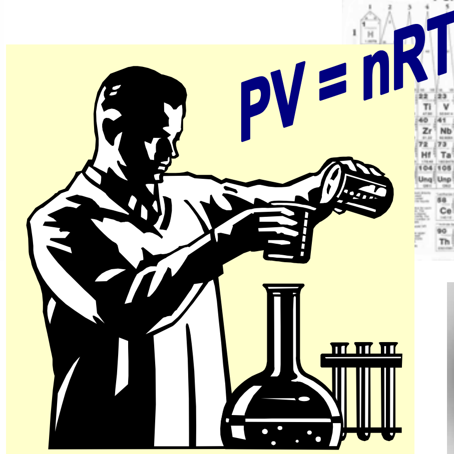
# Project EASE

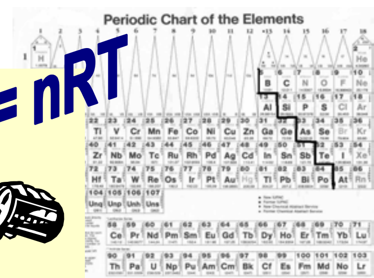
(Effective Alternative Secondary Education)

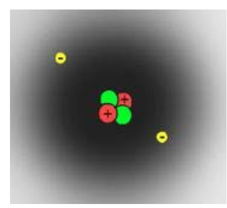
# CHEMISTRY



## MODULE 14

## *The Chemical Bonds*

BUREAU OF SECONDARY EDUCATION
Department of Education
DepEd Complex, Meralco Avenue
Pasig City

# Module 14
# The Chemical Bonds

## What this module is about

Everything around us is made up of matter, from the very simple air that we breathe to the most complex of things that we deal with. You may wonder how we got all these things that make our life easy.

This module is made for you to learn how scientists discovered atoms and how atoms combine and form simple to complex substances. Simple activities are provided to help you understand each lesson. Self-tests are also provided to help you check your understanding and progress in each lesson.

- **Lesson 1 – Ionic and Covalent bonds**
- **Lesson 2 – Polar and Nonpolar molecules**
- **Lesson 3 – Intermolecular Forces of Attraction**

## What you are expected to learn

After going through this module, you are expected to:

1. Differentiate and describe the formation of ionic and covalent bonds.
2. Compare the properties of ionic and covalent compounds.
3. Differentiate between polar and nonpolar molecules
4. Classify molecules as polar or nonpolar.
5. Identify and differentiate the intermolecular forces of attraction.
6. Explain how these forces affect the compound's interaction with other molecules

## How to learn from this module

Here is a simple guide for you in going about this module.

1. Read and follow instructions very carefully.

2. Take the 10-item multiple choice test provided at the start of this module to determine how much you know about the lessons in this module
3. Check your answers against the answer key provided at the last page of the module
4. Perform all the activities provided in each lesson as these will help you have a better understanding of the topic.
5. Take the self-tests at the end of each lesson for you to determine how much you remember about the lesson.
6. Take the 10-item multiple–choice test at the end of the module.

*Good luck and have fun!*

## *What to do before (Pretest)*

**Multiple Choice**. Choose the letter that corresponds to the best answer. Write your choice on a separate paper.

1. It is a type of bond formed when atoms share electrons.
   a. Covalent bond
   b. Ionic bond
   c. Nonpolar bond
   d. Polar bond

2. It refers to the ability of an atom to combine with other atoms
   a. atomic number
   b. atomic mass
   c. ionic number
   d. oxidation number

3. How many electrons are needed in the outer energy levels of most atoms for the atom to be chemically stable?
   a. 2
   b. 4
   c. 6
   d. 8

4. What kind of chemical bond is formed when electrons are gained and lost from atoms?
   a. ionic
   b. covalent
   c. magnetic
   d. metallic

5. How come noble gases do not form compounds readily?
   a. They have no electrons.
   b. They have empty outer energy levels.
   c. They have 7 electrons in their outer energy levels.
   d. Their outer energy levels are completely filled.

6. It is the force that holds the atoms together in a compound.
   a. chemical bond
   b. chemical equation
   c. chemical formula
   d. chemical symbol

7. It refers to a charged atom.
   a. element
   b. ion
   c. compound
   d. molecule

8. Most of the matter around you are in the form of _____.
   a. elements
   b. compounds
   c. atoms
   d. ions

9. A molecule that has oppositely charged ends is called ____.
   a. covalent
   b. ionic
   c. nonpolar
   d. polar

10. Which of the following compounds can form H-bond with water?
   a. $NH_3$
   b. HCl
   c. CaO
   d. $CH_4$

**Key to answers on page 21.**

## Lesson 1. Ionic and Covalent Bonds

### *What you will do*

**Activity 1.1 The Mystery of Table Salt**

1. Get a pinch or two of table salt (NaCl).

2. Place this in a very hot frying pan. Observe what happens after continuously heating it for 15 minutes (be careful not to hold it with your bare hands).

3. Place the same amount of table salt directly on red-hot piece of charcoal. Observe what happens. Is there a difference between the pinch of salt placed in a hot frying fan and the one directly placed on the red-hot charcoal?

4. Sprinkle the same amount of salt in a glassful of water. Observe what happens. What do your observations say regarding the properties of the bond that holds sodium and chlorine together in this compound? Try to compare this with an amount of rubbing alcohol on your skin. Which of the two compounds is being held by a stronger bond?

## *What you will do*

### Activity 1.2  Electrical Conductivity of Solutions

Purpose:

This demonstration provides experimental evidence on the nature of ionic and molecular substances in solution and as a molten ionic solid. Its major purpose is to show that ionic solids conduct electric current both in solution and when fused, whereas molecular solids do not.

Materials:

- Electrical conductivity apparatus (commercial or home-made)
- 6-8 glasses, 50 mL, or 2 oz. wide-mouth bottles
- Solid sodium chloride (table salt), NaCl, and sucrose, $C_{12}H_{22}O_{11}$ (common table sugar)
- Solid silver nitrate, $AgNO_3$
- 2 or 3 small crucibles or brown bottles
- Ring stand, ring, triangle
- 0.1 M solutions of NaCl and sucrose (5.6 g NaCl/100 mL solution and 3.4 g sucrose/100 mL solution)

Safety:

If the conductivity tester used is powered by a 120-V source, then use caution to prevent electric shock.  The substances used are safe and do not require special handling except for the silver nitrate, which is both toxic and caustic.  Handle the silver nitrate with care, particularly when it is fused since it is also an oxidizer. Solutions may be safely disposed by flushing them down the drain with water.  The solid NaCl and sugar may be used for all demonstrations and then disposed in the trash. Keep the crucible with silver nitrate in a dark bottle for later use.  It may be reused many times.

Procedure:

1. Half-fill a 50-mL glass with solid NaCl and a second 50-mL glass with solid sucrose.  Test the electrical conductivity of the solid NaCl with the tester (light bulb remains dark).  Clean the electrodes and then test the solid sucrose (light bulb remains dark).  Record your observation.
2. Half-fill two 50-mL glasses with the two solutions and a third with distilled water. First, test the electrical conductivity of the distilled water.  Then test the electrical conductivity of each of the two solutions in turn.  Record your observation.
3. Half-fill a small crucible with silver nitrate.  Place the crucible on a triangle supported on an iron ring mounted on a ring stand. Heat the crucible with an alcohol lamp until the silver nitrate is melted.  Carefully test the electrical conductivity of the melted silver nitrate.  Record your observation.

## Chemical Bonding: Ionic and Covalent

There are almost 118 elements listed in the periodic table, but there are obviously more substances in nature than 118 pure elements. This is attributed to the capacity of the atoms to combine and react with other atoms to form new compounds. Every compound formed is unique both chemically and physically from its parent atoms.

An example of this is the reaction between the silver-colored metal sodium and the greenish-colored gas chlorine that forms the compound sodium chloride. Sodium metal is a very reactive metal. It reacts violently with water and it produces flame once it gets wet. The element chlorine is a poisonous gas used as a weapon in World War I. When these dangerous substances (sodium metal and chlorine gas) chemically bond together, they form the compound sodium chloride – the common table salt which is so safe that we even eat it everyday!

sodium metal chlorine gas table salt

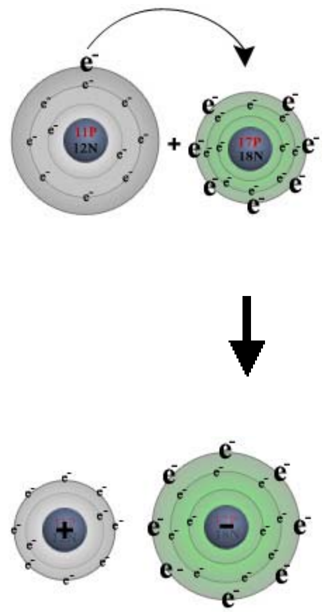
The concept of bonding was originally explained by the American chemist Gilbert Newton Lewis. He proposed that chemical bonds are formed between atoms because electrons from the atoms interact with each other. Lewis had observed that many elements are most stable when they contained eight electrons in their valence shell (outermost shell) like that of the noble gases. This condition is attained by atoms either by transferring or sharing the valence electrons. These two processes of forming the two main types of chemical bonding are *ionic* and *covalent* bonding.

**Ionic bonding** involves the combination of ions of opposite charges. *Cations* are positively-charged atoms which are formed when an atom loses electron/s. The charge depends on the number of electrons the atom lost. On the other hand, atom/s that gain electron/s become negatively charged as the number of electrons exceeds the number of protons and are therefore called *anions*. The compounds formed are called **ionic compounds**.

As shown in the figure on the next page, in ionic bonding electrons are completely transferred from one atom to another. In the process of either losing or gaining negatively charged electrons, the reacting atoms form ions. The oppositely charged ions are attracted to each other by electrostatic forces which are the basis for ionic bond formation. Notice that when sodium loses its one valence electron it gets smaller in size, while chlorine grows larger when it gains an additional valence electron. This is typical of the relative sizes of ions to atoms. Positive ions tend to be smaller than their parent atoms while negative ions tend to be larger than their parent atoms. After the reaction takes place, the charged $Na^+$

and $Cl^-$ ions are held together by electrostatic forces, thus forming an ionic bond. Ionic compounds share many features in common:

- Ionic bonds form between metals and nonmetals.
- In naming simple ionic compounds, the metal is always given first, the nonmetal second (ie. sodium chloride).
- Ionic compounds dissolve easily in water and other polar solvents.
- In solution, ionic compounds easily conduct electricity.
- Ionic compounds tend to form crystalline solids with high melting temperatures.




sodium (on the left) loses one valence electron to chlorine (on the right),

resulting in

a positively charged sodium ion (left) and a negatively charged chlorine ion (right).

the reaction of sodium and chlorine

**What is a Covalent Compound?**

I'm glad you asked! A covalent compound is a compound in which the atoms that are bonded share electrons rather than transfer electrons from one to the other. While ionic compounds are usually formed when metals bond to nonmetals, covalent compounds are formed when two nonmetals bond to each other.

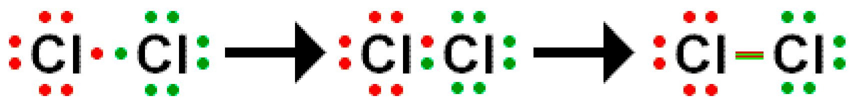
The big question that students frequently ask is: "Why do elements share electrons? After all, wouldn't electrons rather grab electrons outright? That's what happens when ionic compounds are formed."

The reason why nonmetals have to share electrons among themselves has to do with electronegativity. Recall that electronegativity is a measure of how much an element pulls electrons away from other elements it is bonded to. Metals generally have very low electronegativities (they don't want to grab electrons) while nonmetals have high electronegativities (they really want to grab electrons). The reason for this trend is the *octet rule*, which says that all elements want to have the same number of electrons as the nearest noble gas, because noble gases are unusually stable. When metals bond to nonmetals, ionic compounds are formed because the metal atoms don't want electrons and easily give them to nonmetals that want electrons.

It is a different story when two nonmetals bond with each other. Instead of having one element giving electrons to another, we run into a case where we have two elements having roughly the same electronegativity. Thus, neither element can steal electrons from the other. As a result, if either of them is going to be like the nearest noble gas, they'll have to share electrons rather than transfer them.

An example of this case is shown below where two chlorine atoms share their valence electrons. Chlorine atoms have seven electrons each and will be a lot more stable with eight electrons in the outer shell. Single chlorine atoms just do not exist because they come in pairs to share a pair of electrons. The shared pair of electrons make a bond between the atoms. In Lewis structures, the outside electrons are shown with dots and covalent bonds are shown by bars.

This covalent bond between chlorine atoms is one of the most covalent bonds known. Why? A covalent bond is the sharing of a pair of electrons. The two atoms on either side of the bond are exactly the same, so the amount of "pull" of each atom on the electrons is the same, and the electrons are shared equally.



**What are the Properties of Covalent Compounds?**

Covalent compounds have the following properties (keep in mind that these are only general properties, and that there are always exceptions to every rule):

1. Covalent compounds generally have much lower melting and boiling points than ionic compounds.

    As you may recall, ionic compounds have very high melting and boiling points because it takes a lot of energy for all of the + and - charges which make up the crystal to get pulled apart from each other. Essentially, when we have an ionic compound, we need to break all of the ionic bonds in order to make it melt.

    On the other hand, when we have covalent compounds we don't need to break any bonds at all. This is because covalent compounds form distinct molecules, in which the atoms are bound tightly to one another. Unlike in ionic compounds, these molecules don't interact with each other much (except through relatively weak forces called "intermolecular forces"), making them very easy to pull apart from each other. Since they're easy to separate, covalent compounds have low melting and boiling points.

2. Covalent compounds are soft and squishy (compared to ionic compounds).

    The reason for this is similar to the reason that covalent compounds have low melting and boiling points. When you hit an ionic compound with something, it feels very hard because all of the ionic bonds which hold the crystal together tend to make it very inflexible and hard to move. On the other hand, covalent compounds have molecules which easily move around each other because there are no bonds between them. As a result, covalent compounds are more likely to be flexible than hard.

Think of it like this: Ionic compounds are like giant Lego sculptures. If you hit a Lego sculpture with your fist, it feels hard because all of the Legos are stuck very tightly to one another. Covalent compounds are more like those plastic ball pits found in playpens for little kids. While the balls themselves are held together very tightly (just like covalent molecules are held together tightly), the balls aren't really stuck to each other at all. As a result, when little kids jump into the ball pits, they sink in instead of bouncing off.

3. Covalent compounds tend to be more flammable than ionic compounds.

Things burn because they contain carbon and hydrogen atoms that can react to form carbon dioxide and water when heated with oxygen gas (that's the definition of a combustion reaction). Since carbon and hydrogen have very similar electronegativities, they are mostly found together in covalent compounds. As a result, more covalent compounds than ionic compounds are flammable.

There are a couple of exceptions to this rule. The first is with covalent compounds that contain neither carbon nor hydrogen. These tend not to burn, and if they do, they burn by mechanisms other than the classic combustion reaction. The other exception comes with ionic compounds referred to as "organic salts". These organic salts are ionic compounds in which the anion is basically a big covalent molecule containing carbon and hydrogen with just a very small ionic section. As a result, they burn even though they're technically ionic compounds.

4. Covalent compounds don't conduct electricity in water.

Electricity is conducted in water from the movement of ions from one place to the other. These ions are the charge carriers which allow water to conduct electricity. Since there are no ions in a covalent compound, they don't conduct electricity in water.

5. Covalent compounds aren't usually soluble in water.

There's a saying that, "Like dissolves like". This means that compounds tend to dissolve in other compounds that have similar properties (particularly polarity). Since water is a polar solvent and most covalent compounds are fairly nonpolar, many covalent compounds don't dissolve in water. Of course, this is a generalization and not set in stone - there are many covalent compounds that dissolve quite well in water. However, the majority of them don't because of this rule.

## *What you will do*

**Self-Test 1.1**

**Multiple Choice**. Choose the letter that corresponds to the best answer. Write your choice on a separate paper.

1. What is formed when an atom loses electron/s?
   a. anion
   b. cation
   c. neutral atom
   d. none of the above

2. What do you call the electrons which are transferred or shared by an atom?
   a. valence electrons
   b. inner- shelled electrons
   c. excited electrons
   d. electrons in the ground state

3. Which is a stronger bond?
   a. covalent
   b. ionic
   c. single bond
   d. double bond

4. What rule or law is followed by atoms in attaining eight electrons in the outermost shell?
   a. Law of definite proportion
   b. Law of conservation of matter
   c. Law of Multiple proportion
   d. Octet rule

5. Which compound has the higher boiling point/melting point?
   a. alcohol
   b. table salt
   c. oil
   d. wax

**Key to answers on page 22.**

*If you scored 3 or more out of 5 –CONGRATULATIONS!*

*You may now proceed to the next lesson. If you scored below 3, you need to go back and read the lesson again!*

## Lesson 2. Polar and Nonpolar Bonds

### *What you will do*

**Activity 2.1**

Why is it that oil and water do not mix?

In your school, do the boys and girls mix? Probably not and that is not because the girls don't like the boys (usually), or because the boys don't like the girls. It's just that young girls like spending time with other young girls more than with young boys. Boys and girls tend to like to do different things and talk about different things when they are young.

Imagine you are a girl going out of the classroom for recess. You look around. You see a boy or two, and a girl or two. Naturally you go over to the girls to talk. Later, another girl comes out. She, too, goes over to the girls to talk. After a while, after every child is outside the classroom you have batches of boys and batches of girls all over the place, but not many areas where boys and girls are together.

The same thing happens with oil and water. Oil is made of molecules (which are like very small, sticky rubber balls). So is water. But water molecules are not the same as oil molecules. Most importantly, while all molecules like to stick to each other, oil molecules like to stick to other oil molecules more than they like to stick to water molecules. Same with the water: water molecules like to stick to oil molecules, but they like to stick to other water molecules even more.

So, if you pour some water into oil, the water molecules coming in see oil molecules and (at first only a few) water molecules. Naturally they prefer to join the other water molecules. Thus, you get particles of water and particles of oil, but nowhere do you find the two kinds of molecules mixed up together. It turns out that oil is lighter than water, so the particles of oil tend to float up on top of the water, and join together into a big oil slick. If you get enough oil, you get a thick layer of it on top of the water. You can sort of mix up the oil and water by shaking the jar, but you will only break up the layer into smaller particles, and they will come back together again quickly because of the lightness of oil. Now, if you were on the Space Shuttle, the particles wouldn't form a thick layer, they'd just stay all jumbled up. The oil and water wouldn't mix!

You can force oil and water to mix. What you need is a type of molecule that both the water and the oil like to stick to more than they like to stick to each other. Such a molecule is called an emulsifier, or, more simply, a soap. When you get a soap molecule, the water and the oil stick to it, and then the oil and water mix.

A lot of things can be used as emulsifiers. Soap is one of them. If you get grease on your hands, plain water won't wash it off -- the grease won't mix with the water and come off. But if you put soap on your hands, the grease will mix with the water and come off.

An egg yolk is another emulsifier. If you mix salad oil and vinegar (which is mostly water), and then put some eggs in, the oil and vinegar will mix -- you get mayonnaise. You can make mayonnaise with oil, vinegar, and eggs in your blender at home.

Water is a polar molecule which means that it has a negatively charged end and a positively charged end. So water molecules attract each other. They also attract other polar molecules. Oil is NOT a polar molecule-it doesn't have a separation of charge. So water and oil aren't attracted to each other. Just remember-***like dissolves like.***

## Polar Covalent Compounds

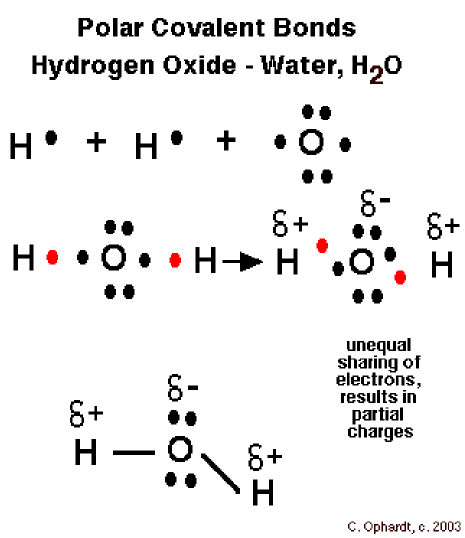
Bonding between nonmetals consists of two electrons shared between two atoms. In covalent bonding, the two electrons shared by the atoms are attracted to the nucleus of both atoms. Neither atom completely loses or gains electrons as in ionic bonding.

There are two types of covalent bonding:

1. Nonpolar bonding with an equal sharing of electrons.
2. Polar bonding with an unequal sharing of electrons. The number of shared electrons depends on the number of electrons needed to complete the octet.

**POLAR BONDING** is the result when two different nonmetals unequally share electrons between them. One well known exception to the identical atom rule is the combination of carbon and hydrogen in all organic compounds.

The nonmetal closer to fluorine in the Periodic Table has a greater tendency to keep its own electron and also draw away the other atom's electron. It is NOT completely successful. As a result only partial charges are established. One atom becomes partially positive since it has lost control of its electron most of the time. The other atom becomes partially negative since it gains electron most of the time.

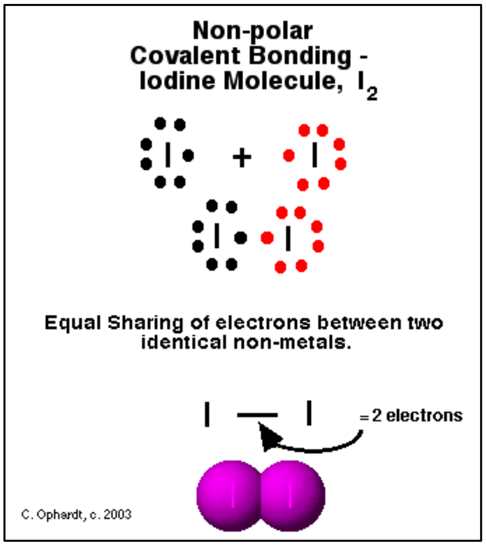



Water, the most universal compound on earth, has the property of being a polar molecule. Thus, the physical and chemical properties of the compound are fairly unique.

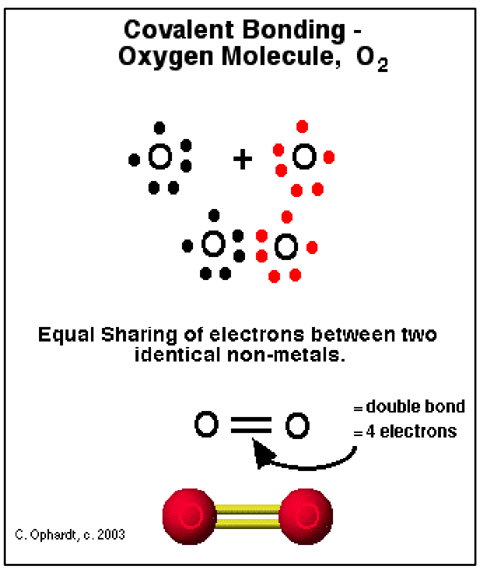
Hydrogen Oxide or water forms a polar covalent molecule. The graphic on the left shows that oxygen has 6 electrons in the outer shell. Hydrogen has one electron in its outer energy shell. Since 8 electrons are needed for an octet, they share the electrons.

However, oxygen gets an unequal share of the two electrons from both hydrogen atoms. Again, the electrons are still shared (not transferred as in ionic bonding), and the sharing is unequal. The electrons spend more of the time closer to oxygen. As a result, the oxygen acquires a "partial" negative charge. At the same time, since hydrogen loses the electron most - but not all of the time, it acquires a "partial" charge. The partial charge is denoted with a small Greek symbol for delta. Water is a polar molecule with positive charges on one side and negative charges on the other.

**NONPOLAR BONDING** results when two **identical nonmetals equally share** electrons between them. One well known exception to the identical atom rule is the combination of carbon and hydrogen in all organic compounds. A nonpolar molecule is one where the electrons are distributed more symmetrically. Thus it does not have an abundance of charges at the opposite sides. The charges all cancel each other out.




Iodine forms a diatomic nonpolar covalent molecule. The graphic on the top left shows that iodine has 7 electrons in the outer shell. Since 8 electrons are needed for an octet, two iodine atoms EQUALLY share 2 electrons.

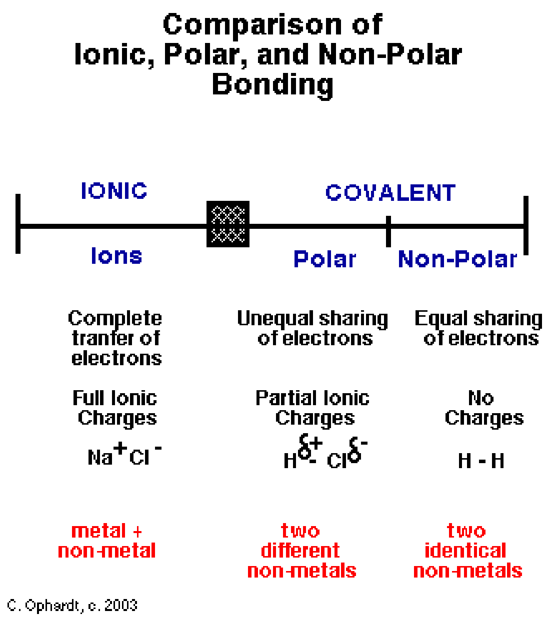


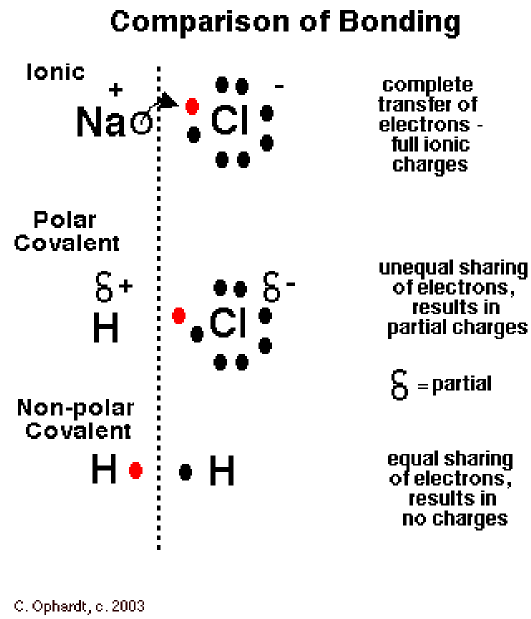

Molecules of oxygen, which are present in about 20% concentration in air, are also covalent molecules. (See the graphic at the left). There are 6 electrons in the outer shell, therefore, 2 electrons are needed to complete the octet. The two oxygen atoms share a total of four electrons in two separate bonds called double bonds. The two oxygen atoms **equally share** the four electrons.

**Comparison of Ionic, Polar, and Nonpolar Bonds**

Whereas nonpolar bonding involves the equal sharing of electrons between identical nonmetal atoms, POLAR BONDING is the unequal sharing of electrons between two different nonmetal atoms. A proper understanding of polar bonding is gained by viewing the types of bonding on a continuum as in the diagram below. Ionic bonding is on one extreme with a complete transfer of electrons forming charged ions.

Nonpolar covalent bonding with equal sharing of electrons is at the other extreme. Somewhere in the middle but favoring the covalent side is polar bonding with unequal sharing of electrons and partial but incomplete transfer of electrons.




**Rule:**

Polar molecules will mix to form solutions and nonpolar molecules will also mix to form solutions, but a polar and nonpolar combination will not form a solution.

Water is a polar molecule and oil is a nonpolar molecule. Thus they won't form a solution. On the other hand, since alcohol is a polar molecule, it will form a solution with water.

Before going on, try answering the following:

## *What you will do*

**Self-Test 2.1**

**Multiple Choice**. Choose the letter that corresponds to the best answer. Write your choice on a separate paper.

1. Which of the following pairs will form polar covalent bond?
   a. Cl and Cl
   b. H and Cl
   c. H and H

2. Where will the partial positive end be found on the bond that will be formed in number 1?
   a. H
   b. Cl

3. Which of the following pairs will be soluble with one another?
    a. $Br_2$ and $CCl_4$
    b. Oil and water
    c. $H_2O$ and $CCl_4$

For numbers 4 and 5, refer to the following choices:
    a. polar covalent
    b. nonpolar covalent
    c. none of the above

4. What kind of bond is formed by two identical atoms?
5. It is the kind of bond formed by atoms with different electronegativities.

**Key to answers on page 22.**

*If you scored 3 or more out of 5 –CONGRATULATIONS!*

*You may now proceed to the next lesson. If you scored below 3, you need to go back and read the lesson again!*

## Lesson 3. Intermolecular Forces of Attraction

### *What you will do*

**Activity 3.1**

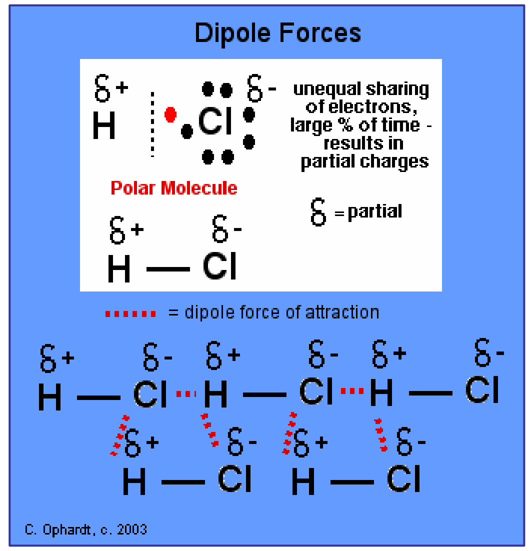
How would you explain the difference in the boiling points of water and rubbing alcohol?

The physical properties of substances like melting point, boiling point, vapor pressure, evaporation, viscosity, surface tension, and solubility are related to the strength of attractive forces between molecules. These attractive forces are called **Intermolecular Forces**. This is the force responsible for why molecules stick together. The amount of "stick togetherness" is important in the interpretation of the various properties listed above.

The three types of intermolecular forces are **Dispersion Forces, Dipole-dipole interactions** and **Hydrogen bonding**. Dispersion forces, also known as London Forces, (named after Fritz London who first described these forces theoretically 1930) is considered as the weakest intermolecular force. Most of the intermolecular forces are identical to bonding between atoms in a single molecule.

## Dipole Forces

Polar covalent molecules are sometimes described as "dipoles", meaning that the molecule has two "poles". One end (pole) of the molecule has a partial positive charge while the other end has a partial negative charge. The molecules will orientate themselves so that the opposite charges attract and the principle operates effectively.

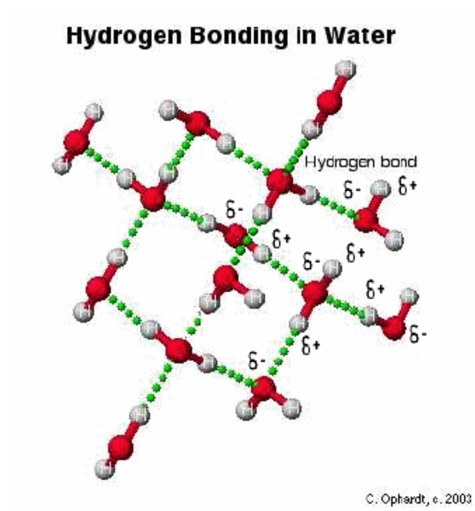



In the example on the right, hydrochloric acid is a polar molecule with the partial positive charge on the hydrogen and the partial negative charge on the chlorine. A network of partial + and - charges attract molecules to each other.

Dipole-dipole interactions are stronger intermolecular forces than Dispersion forces. They occur between molecules that have permanent net dipoles (polar molecules). For example, dipole-dipole interactions occur between $SCl_2$ molecules, $PCl_3$ molecules and $CH_3Cl$ molecules. If the permanent net dipole within the polar molecules results from a covalent bond between a hydrogen atom and either fluorine, oxygen or nitrogen, the resulting intermolecular force is referred to as a hydrogen bond (see below). The partial positive charge on one molecule is electrostatically attracted to the partial negative charge on a neighboring molecule.

## Hydrogen Bonding

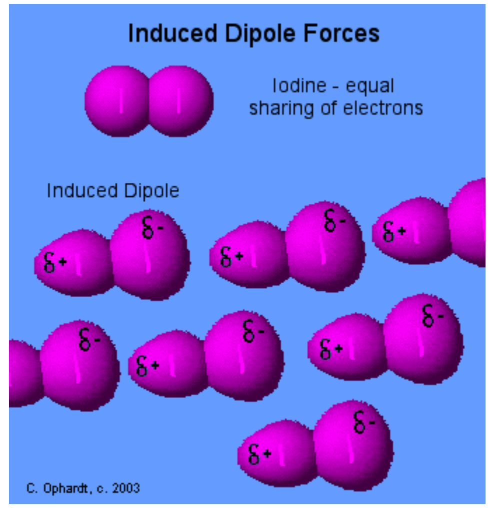
The hydrogen bond is really a special case of dipole forces. A hydrogen bond is the attractive force between the hydrogen attached to an electronegative atom of one molecule and an electronegative atom of a different molecule. Usually the electronegative atom is oxygen, nitrogen, or fluorine. To recognize the possibility of hydrogen bonding, examine the Lewis structure of the molecule. The electronegative atom must have one or more unshared electron pairs as in the case of oxygen and nitrogen, and has a negative partial charge. The hydrogen, which has a partial positive charge, tries to find another atom of oxygen or nitrogen with excess electrons to share and is attracted to the partial negative charge. This forms the basis for the hydrogen bond.




In the figure on the left, the hydrogen is partially positive and is attracted to the partially negative charge on the oxygen. Because oxygen has two lone pairs, two different hydrogen bonds can be made to each oxygen. H bonds are stronger intermolecular forces than either dispersion forces or dipole-dipole interactions.

This is a very specific bond as indicated. Some combinations which are not hydrogen bonds include hydrogen to another hydrogen or hydrogen to a carbon.

### Induced Dipole Forces (Dispersion Forces)




Forces between essentially nonpolar molecules are the weakest of all intermolecular forces. "Temporary dipoles" are formed by the shifting of electron clouds within molecules. These temporary dipoles attract or repel the electron clouds of nearby nonpolar molecules.

The temporary dipoles may exist for only a fraction of a second but a force of attraction also exist for that fraction of time. The strength of induced dipole forces depends on how easily electron clouds can be distorted. Large atoms or molecules with many electrons far removed from the nucleus are more easily distorted.

Induced dipole or dispersion forces are very weak forces of attraction between molecules because this interaction is the result of the following:

- momentary dipoles occurring due to uneven electron distributions in neighbouring molecules as they approach one another
- the weak residual attraction of the nuclei in one molecule for the electrons in a neighbouring molecule.

The more electrons that are present in the molecule, the stronger the dispersion forces will be. Dispersion forces are the only type of intermolecular force operating between nonpolar molecules. For example, dispersion forces operate between hydrogen ($H_2$) molecules, chlorine ($Cl_2$) molecules, carbon dioxide ($CO_2$) molecules, dinitrogen tetroxide ($N_2O_4$) molecules and methane ($CH_4$) molecules.

### Relative Strength of Intermolecular Forces

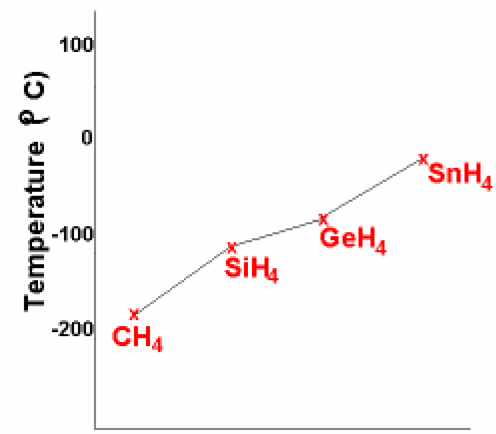
Intermolecular forces (dispersion forces, dipole-dipole interactions and hydrogen bonds) are much weaker than intramolecular forces (covalent bonds, ionic bonds or metallic bonds). Dispersion forces are the weakest intermolecular force (one hundredth-one thousandth the strength of a covalent bond). On the other hand, hydrogen bonds are the strongest intermolecular force (about one-tenth the strength of a covalent bond).

(dispersion forces < dipole-dipole interactions < hydrogen bonds)

**Effect of Intermolecular Forces on Melting and Boiling Points of Molecular Covalent Substances**

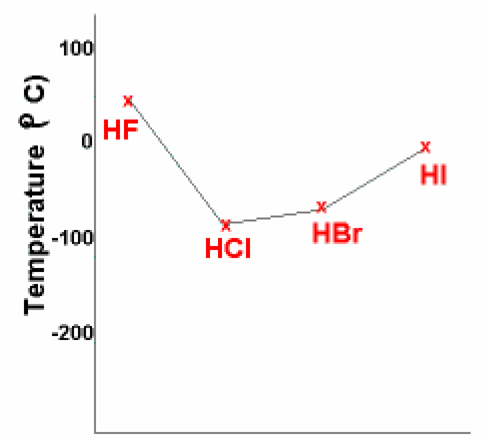
Since melting or boiling results from a progressive weakening of the attractive forces between the covalent molecules, the stronger the intermolecular force is, the more energy is required to melt the solid or boil the liquid.

If only dispersion forces are present, then the more electrons the molecule has (and consequently the more mass it has) the stronger the dispersion forces will be, and the higher the melting and boiling points will be. Consider the hydrides of Group IV, all of which are nonpolar molecules; thus, only dispersion forces act between the molecules. $CH_4$ (molecular mass ~ 16), $SiH_4$ (molecular mass ~ 32), $GeH_4$ (molecular mass ~ 77) and $SnH_4$ (molecular mass ~ 123) can all be considered nonpolar covalent molecules.




As the mass of the molecules increases, so does the strength of the dispersion force acting between the molecules. This implies that more energy is required to weaken the attraction between the molecules resulting in higher boiling points.

If a covalent molecule has a permanent net dipole then the force of attraction between these molecules will be stronger than if only dispersion forces were present between the molecules. As a consequence, this substance will have a higher melting or boiling point than similar molecules that are nonpolar in nature. Consider the boiling points of the hydrides of Group VII elements. All of the molecules HF (molecular mass ~ 20), HCl (molecular mass ~ 37), HBr (molecular mass ~ 81) and HI (molecular mass ~ 128) are polar, the hydrogen atom having a partial positive charge (H) and the halogen atom having a partial negative charge(F, Cl, Br, I).




As a consequence, the stronger dipole-interactions acting between the hydride molecules of Group VII elements results in higher boiling points than for the hydrides of Group IV elements as seen in the previous page.

Usually, as the <u>molecular mass</u> increases, the boiling point of the hydrides increases. HF is an exception to this rule because of the stronger force of attraction between HF molecules resulting from hydrogen bonds acting between the HF molecules. Weaker dipole-dipole interactions act between the molecules of HCl, HBr and HI. So HF has a higher boiling point than the other molecules in this series.

**Effect of Intermolecular Forces on Solubility**

In general like dissolves like. Specifically

- *nonpolar solutes dissolve in nonpolar solvents.* Paraffin wax ($C_{30}H_{62}$) is a nonpolar solute that will dissolve in nonpolar solvents like oil, hexane ($C_6H_{14}$) or carbon tetrachloride ($CCl_4$). Paraffin wax will NOT dissolve in polar solvents such as water ($H_2O$) or ethanol (ethyl alcohol, $C_2H_5OH$).

- polar solutes such as glucose ($C_6H_{12}O_6$) will dissolve in polar solvents such as water ($H_2O$) or ethanol (ethyl alcohol, $C_2H_5OH$). This is so because the partially positively charged atom of the solute molecule is attracted to the partially negatively charged atom of the solvent molecule, and the partially negatively charged atom of the solute molecule is attracted to the partially positively charged atom of the solvent molecule. Glucose will NOT dissolve in nonpolar solvents such as oil, hexane ($C_6H_{14}$) or carbon tetrachloride ($CCl_4$).

- Ionic solutes such as sodium chloride (NaCl) will generally dissolve in polar solvents but not in nonpolar solvents, since the positive ion is attracted to the partially negatively charged atom in the polar solvent molecule, and the negative ion of the solute is attracted to the partially positively charged atom on the solvent molecule.

Before going on, try answering the following:

## *What you will do*

**Self-Test 3.1**

**Multiple Choice**. Choose the letter that corresponds to the best answer. Write your choice on a separate paper.

1. The boiling point of $CH_4$ is much lower than that of HF. This is explained by the
    a. hydrogen bonding in HF
    b. ion-dipole interaction in $CH_4$
    c. polarity of $CH_4$

2. What causes the high surface tension of water?
    a. hydrogen bonding
    b. dipole-dipole
    c. dispersion forces

3. Which is most likely to have the strongest intermolecular forces?
    a. solid
    b. liquid
    c. gas

For nos. 4 and 5, refer to the following choices:

a. hydrogen bond
b. dipole-dipole
c. dispersion forces

What intermolecular force of attraction is present in each of the following?

4. water
5. carbon dioxide

**Key to answers on page 22.**

*If you scored 3 or more out of 5 –CONGRATULATIONS!*
*If you scored below 3, you need to go back and read the lesson again!*

## *Let's Summarize*

1. There are two types of bond that can be produced when atoms combine. These are:

   a. Ionic bond

   This kind of bond is formed by the electrostatic attraction between a cation (metal) and an anion (nonmetal). The compound formed is called ionic compound. Ionic compounds have higher melting and boiling points compared to covalent compounds. They easily ionize in water.

   b. Covalent bond

   The bond is formed when nonmetals share their valence electrons (electrons in the outermost shell) with another nonmetal. The compound formed is called covalent compound and it may be diatomic or polyatomic in nature. Generally, covalent bonds are weaker than ionic compounds, and they have lower boiling and melting points, too.

2. Classification of Covalent Bonds

   a. Polar covalent bond is formed when there is unsymmetrical sharing of electron cloud. It is formed by nonmetals of different electronegativity. This bond is characterized by the presence of partial positive and negative ends. The partial positive end is found on the atom with the lower electronegativity while partial negative end is on the more electronegative atom.
   b. Nonpolar covalent bond is characterized by the symmetrical distribution of electron cloud around the combining atoms. Identical atoms form this kind of

bond.

   c. The rule of thumb that *"like dissolves like*" is best explained by this concept on polarity.

3. Three types of force can operate between covalent molecules:

   a. Dispersion Forces is also known as London Forces (named after Fritz London who first described these forces theoretically in 1930) or as Weak Intermolecular Forces or as van der Waal's Forces (named after the person who contributed to our understanding of non-ideal gas behavior).
   b. Dipole-dipole interactions
   c. Hydrogen bonds

4. Relative strength of Intermolecular Forces:

   a. Intermolecular forces (dispersion forces, dipole-dipole interactions and hydrogen bonds) are much weaker than intramolecular forces (covalent bonds, ionic bonds or metallic bonds)
   b. Dispersion forces are the weakest intermolecular force (one hundredth-one thousandth the strength of a covalent bond), hydrogen bonds are the strongest intermolecular force (about one-tenth the strength of a covalent bond).
   c. dispersion forces < dipole-dipole interactions < hydrogen bonds

## *Posttest*

**Multiple Choice**. Choose the letter that corresponds to the best answer. Write your choice on a separate paper.

1. Solute is most likely to be highly soluble in a solvent if the solute is _____ and the solvent is ______.
   a. ionic, polar
   b. nonpolar, ionic
   c. polar, polar
   d. nonpolar, ionic

2. Molecular iodine would be most soluble in
   a. Water
   b. Carbon tetrachloride
   c. Ethanol and water

3. The concept of "like dissolves like" is illustrated by which of the following?
   a. NaCl (s) is more soluble in $CCl_4$ than in water.
   b. $I_2$ (s) is more soluble in $CCl_4$ than in water.
   c. $CCl_4$ is soluble in water

4. Which one of the following molecules would be most polar?
    a. $H_2$
    b. HF
    c. HCl

5. Which of the following bond types is the strongest?
    a. Polar
    b. Nonpolar
    c. Ionic

For numbers 6-8, refer to the following choices:
    a. covalent bond
    b. ionic bond

6. The bond formed by $Ca^{2+}$ and $O^{2+}$
7. The bond formed by two oxygen atoms
8. The bond present in table salt

For numbers 9 to 10, refer to the following:
    a. $BaCl_2$
    b. $CO_2$
    c. $NH_3$

9. Which of the given compounds has the highest boiling point?
10. Which one will not dissolve in a polar solvent?

**Key to answers on page 22.**

*If you scored 7 or more out of 10 – Congratulations! You may now proceed to the next module. If you scored below 7, you need to go back and read the module again!*

## *Key to Answers*

**Pretest**

| | |
|---|---|
| 1. a | 6. a |
| 2. d | 7. b |
| 3. d | 8. b |
| 4. a | 9. d |
| 5. d | 10. a |

| Lesson 1 | Lesson 2 | Lesson 3 |
|---|---|---|
| **Self-Test 1.1** | **Self-Test 2.1** | **Self-Test 3.1** |
| 1. b | 1. b | 1. a |
| 2. a | 2. a | 2. a |
| 3. b | 3. a | 3. a |
| 4. c | 4. b | 4. a |
| 5. b | 5. a | 5. c |

**Posttest**

| | |
|---|---|
| 1. c | 6. b |
| 2. b | 7. a |
| 3. b | 8. b |
| 4. b | 9. a |
| 5. c | 10. b |

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## *References*


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