# Chemical Bonding Script:

In this video we will learn about bonding. One of the distinguishing features of how chemists and molecular scientists describe the world focuses on how atoms and molecules interact with each other through bonding.

Before starting our talk about bonding, we should go over some definitions. The first is molecule, a molecule is two or more bonded atoms. A molecular element is two or more bonded atoms of the same type, like a pure sample of silver, or a molecule of oxygen gas. A molecular compound is two or more bonded atoms of different types, like salt or the caffeine in coffee. A pure substance is two or more molecules or atoms of the same type. This could be either a canister of carbon dioxide gas, or a vial full of liquid mercury.

Now that we have some of the terminology down we can move on to bonding concepts. Bonding between atoms is due to the forces described by Coulombs law. For our purposes it is enough to know that because protons and electrons are charged particles they are subject to attractive and repulsive electronic forces. The charged particle interactions in an atom are as follows: The interaction between two electrons is repulsive, the interaction between two nuclei is repulsive, and the interactions between electrons and nuclei is attractive. If the attraction between two atoms is greater than the repulsion a bond is formed.

There are different classes of chemical bonding, in this presentation we will be talking about ionic, covalent and polar covalent bonding.

An Ionic bond is generally formed between metals and non-metals. When metal and nonmetal atoms approach each other, an electron transfer takes place. The metal transfers one or more electrons to the nonmetal so both atoms reach the nearest noble gas configuration. This example shows a sodium atom transferring an electron to a chlorine atom. The resulting charged ions are then attracted to each other due to electrostatic forces. The positively charged sodium cation is attracted to the negatively charged chlorine anion forming the ionic solid sodium chloride, commonly known as table salt. Simply put, Ionic bonding is the electrostatic attraction between positive and negative ions.

A covalent bond is equal electron sharing between bonded, nonmetal atoms. Instead of transferring electrons between bonding atoms, bonded nonmetals will share valence electrons to achieve a noble gas configuration. In a covalent bond the electrons are shared equally between the bonded atoms. Sharing electrons allows two bonded hydrogen atoms to attain the nearest noble gas configuration of helium. The area of highest electron density will be between the two bonded hydrogen atoms, each hydrogen now has access to two electrons.

Polar covalent bonding is the unequal sharing of electrons between atoms. Molecules that have unequal electron sharing have a partial charge separation on the bonded atoms. The atom in a molecule that is more electronegative will have a larger electron density and a slight net negative charge, termed delta negative.  The atom in a molecule that is less electronegative will have a smaller electron density and a slight net positive charge, termed delta positive.  This hydrogen fluoride molecule has a higher electron density centered on the fluorine atom because fluorine is more electronegative than hydrogen.  The fluorine is delta negative, and the hydrogen is delta positive.

The electronegativity difference ($\Delta$EN) between bonded atoms can be used to identify which type of bond exists in a molecule.  A purely covalent bond, where electrons are shared equally, will have no charge on the bonded atoms, and an electronegativity difference of 0 to 0.4 $\Delta$EN.  A polar covalent bond, where electrons are shared unequally, will have a partial charge on the bonded atoms, and an electronegativity difference of 0.4 to 2.0 $\Delta$EN. An ionic bond, where electrons are transferred, will have a full charge on the bonded ions, and an electronegativity difference of greater than 2.0 $\Delta$EN.

This figure has electrostatic diagrams of oxygen, hydrogen chloride and potassium bromide. An electrostatic diagram is a three dimensional picture of a molecule that demonstrates the charge distribution.  Red indicates an area of high electron density, blue indicates an area of low electron density, and white indicates no charge.  The covalently bonded oxygen has no charge separation, hydrogen chloride has a partial charge separation and potassium chloride has a full charge on each ion.

Before we continue, we need to expand our definition of valence electrons.  We know that valence electrons are the electrons in the outermost shell of an atom.  Most importantly, valence electrons are the electrons primarily involved in chemical bonding between atoms.  The number of valence electrons an element has determines the chemical properties of that element.  Elements in a group have the same number of valence electrons, this is why elements in a group have similar chemistry.  The valence electrons of a main group element are located in the outermost shell.  The valence electrons of a transition metal are located in the outermost d orbitals, as well as the outermost shell. Remember that the outermost shell is the highest value of n the principle quantum number.

We are ultimately interested in valence electrons because they are responsible for much of the chemical behavior of atoms.  We use Lewis structures to represent the valence electrons of a main group element using dots surrounding the chemical symbol.

Lewis theory is the idea that chemical bonding is simply the attainment of a stable electron configuration through the sharing or transfer of electrons between bonded atoms.  Lewis theory uses the octet rule to predict bonding.  An octet is a full outer shell containing eight electrons.  Lewis structures can be used to demonstrate bonding in molecules by showing atoms in a molecule sharing electrons to attain a full octet.  This figure shows the Lewis structures for the first ten elements of the periodic table.  In the first row of the periodic table helium, has a full valence shell with two electrons.  You can see that neon has a full octet.  This is why the noble gases are generally unreactive, they have a full octet of electrons.

The octet rule can be used to predict how atoms bond based on the premise that, a stable electron configuration can be attained with eight electrons in the outermost shell.

Bonding atoms will transfer or share electrons to satisfy the octet rule, this way each atom will have access to eight electrons in its outermost shell. It is important to realize that the octet rule only works for the second row of the periodic table. Elements beyond the second row can access d or f orbitals, elements are larger and have more room for bonding. This means that elements beyond the second row (after Neon) can have expanded octets, this is referred to as Hyper-coordination. Despite its limitations, the octet rule remains a useful tool for predicting bonding in molecules.

The last portion of this video will be dedicated to Lewis structures, we use Lewis structures to demonstrate bonding between atoms. You can see in this figure how the filling of valence shells for bonded atoms is displayed using Lewis structures. There are detailed example videos on drawing Lewis structures to watch after this video. For now, become accustomed to the methodology and terminology used for drawing Lewis structures.

There is a basic method for drawing molecular Lewis structures.

1) Determine the total number of valence electrons for the entire molecule by adding the amount of valence electrons for each molecule, for anions add an extra electron, for cations subtract an electron.
2) Write the correct structure for the molecule, use a line to connect the bonding atoms, each line represents two electrons. Subtract bonding electrons from total electrons.
    - Hydrogen will always be terminal.
    - Generally the least electronegative atom will be in the center.
3) Distribute remaining electrons (as pairs) first to the terminal atoms, then to the central atom. Subtract from your total as you fill octets.
    1) Hydrogen has a full shell with two electrons. Do not assign electrons (other than bonding) to hydrogen.
4) If there are any atoms without a full octet, move electrons to form double or triple bonds, use arrows to indicate electron movement. Redraw your structure.

When you begin to draw Lewis structures you will quickly notice that there is often more than one structure that can be drawn. To choose the most correct structure assign formal charge to each atom in the structure. Formal charge is the charge on an atom in a Lewis structure, assuming all electrons are shared equally. You can calculate formal charge with the following formula: Valence electrons, minus one half of the bonding electrons, minus lone electrons equals formal charge. The structure with the lowest possible charge indicates the most energetically favorable structure.

It is possible to have more than one correct Lewis structure, in this case both structures are used. Resonance structures is the term used when more than one Lewis structure is allowed for a molecule. Drawing both structures with a double headed arrow between them is the correct way to indicate resonance.

We have now covered the basics of bonding between atoms. You should be able to define molecule, molecular element, molecular compound, and pure substance. You should be able to explain ionic, covalent and polar covalent bonding as well as identify the type of bond in a molecule using the electronegativity difference. You should have a good conceptual understanding of valence electrons, Lewis theory, and the octet rule, and have learned the methods and terminology for drawing molecular Lewis structures. Be sure to watch the videos detailing how to draw Lewis structures before attempting the problem set for this section.