Chemical Bonding

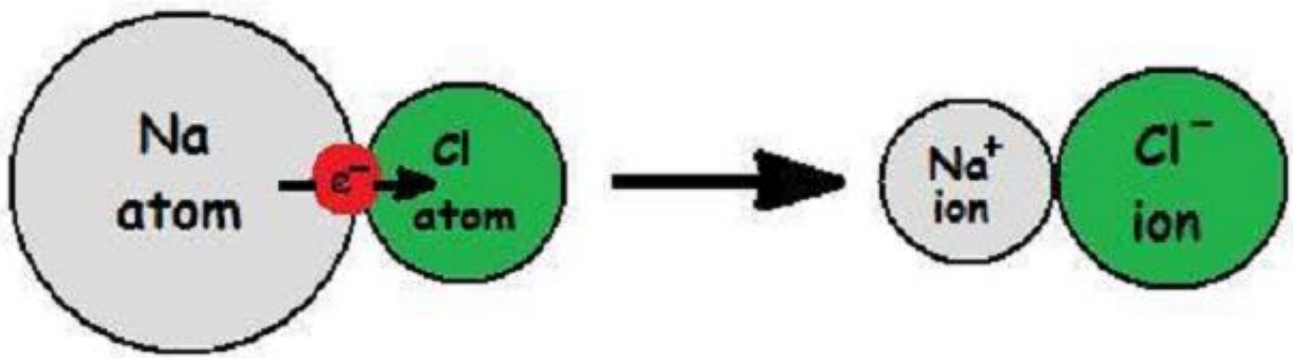
Ionic Bond

A bond between ions resulting from the transfer of electrons from one of the bonding atoms to the other and the resulting electrostatic attraction between the ions.

Ionic Bonding

- Molecular collisions between atoms that tend to lose electrons (metals) and atoms that tend to gain electrons (non=-metals) are sometime sufficient to remove electrons from the metal atom and add them to the non-metal atom.
- This transference of electrons from metals to non-metals forms positive and negative ions which in turn, attract each other due to opposite charges.
- The compounds formed by this electrostatic attraction are said to be ionically bonded.

- Ionic bonds are formed by a transfer of electrons from metal atoms to non-metal atoms with the resulting electrostatic attraction holding the ions together.
- The octet rule is an expression of the tendency for atoms gain or lose the appropriate number of electrons so that the resulting ion has either completely filled or completely empty outer energy levels.



During a collision between sodium and chlorine atoms, an electron is transferred.

Covalent Bonding

Introduction

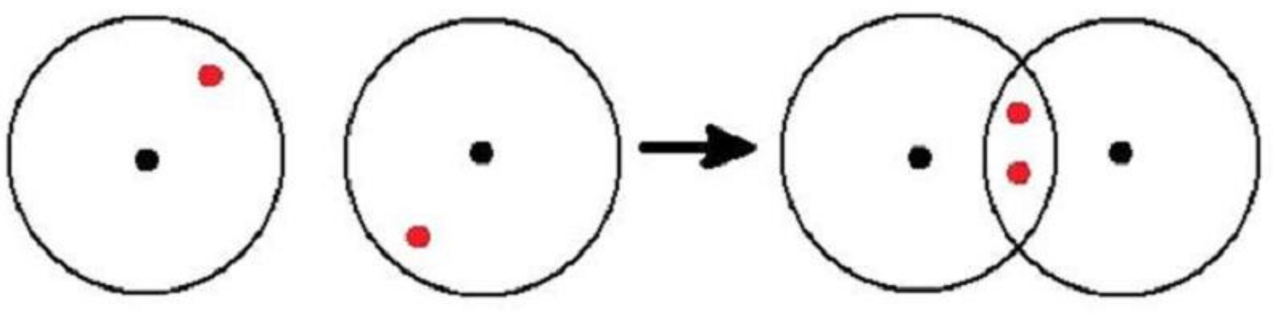
- In ionic bonding, electrons leave metallic atoms and enter non-metallic atoms.
- This complete transfer of electrons changes both of the atoms into ions.
- Often, however, two atoms combine in a way that no complete transfer of electrons occurs.
- Instead, electrons are held in overlapping orbitals of the two atoms, so that the atoms are sharing the electrons.

- The shared electrons occupy the valence orbitals of both atoms at the same time.
- The nuclei of both atoms are attracted to this shared pair of electrons and the atoms are held together by this attractive force.
- The attractive force produced by sharing electrons is called a covalent bond.

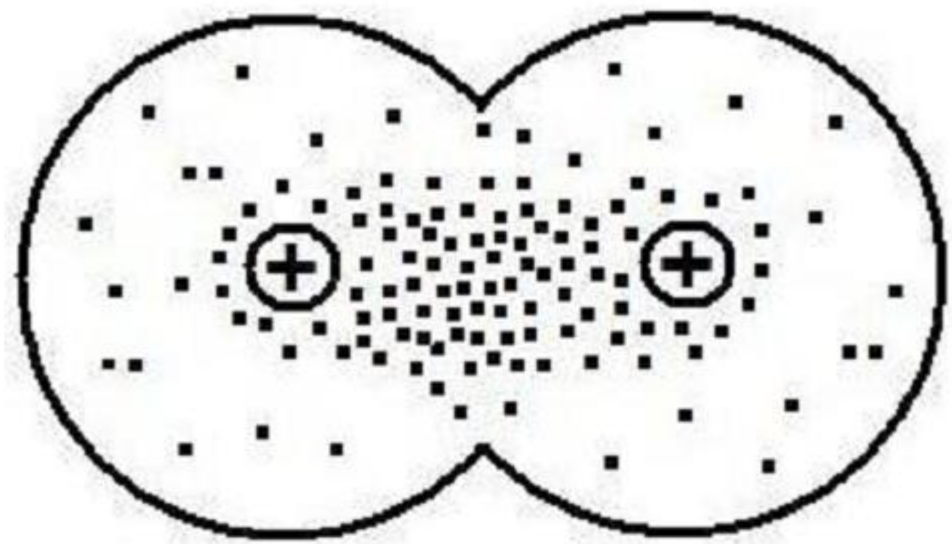
Sharing Electrons

- In covalent bonding, the atoms acquire a stable octet of electrons by sharing electrons.
- The covalent bonding process produces molecular substances as opposed to the lattice structures of ionic bonding.
- The covalent bond, in general, is much stronger than ionic bonds and there are far more covalently bonded substances than ionic substances.

- The diatomic hydrogen molecule, H2, is one of the many molecules that are covalently bonded.
- Each hydrogen atom has a 1s electron cloud containing one electron.
- These 1s electron clouds overlap and produce a common volume which the two electrons occupy.



The 1s orbitals of two hydrogen atoms overlap and share the two valence electrons.



Simulated probability pattern for the overlapped orbitals in H2.

- The diatomic fluorine molecule is also a covalently bonded atom. In the case of fluorine atoms, the atoms have filled 1s orbitals, filled 2s orbitals, and two of the three 2p orbitals are full.
- Each atom has a half-filled 2p orbital that is available to be overlapped.

A covalent bond between atoms with an electron dot formula where the shared pair of electrons are the bonding electrons or with the bond represented by a dash.

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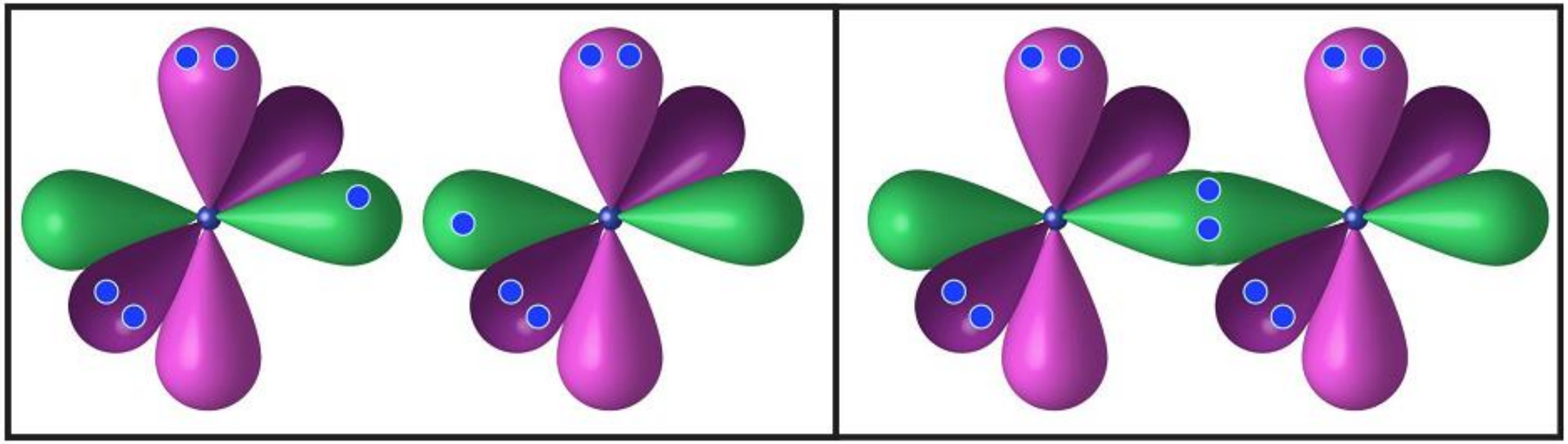
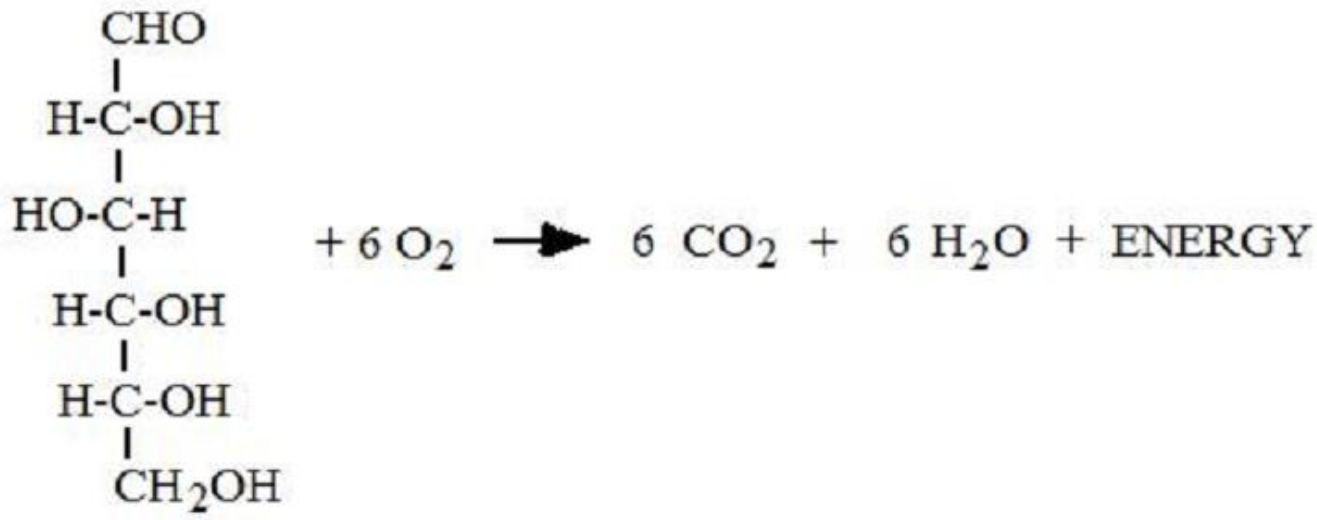


Showing the 2p orbitals of fluorine with two orbitals full and one orbital half-full and then showing the half-filled orbital of each atom overlapping.

Molecular Stability

- Compounds with high bond strengths are difficult to break up and therefore are stable compounds.
- When stable compounds are formed, large amounts of energy are given off so these molecules are in a relatively low energy state.
- In order to break the molecule apart, all the energy that was given off must be put back in.
- Low energy state bonds are stable and have high bond strength.

- Molecules with high bond energy have weak bonds.
- They did not release much energy when they formed and so not much energy is needed to the break the molecules back apart.
- High bond energy means low bond strength.

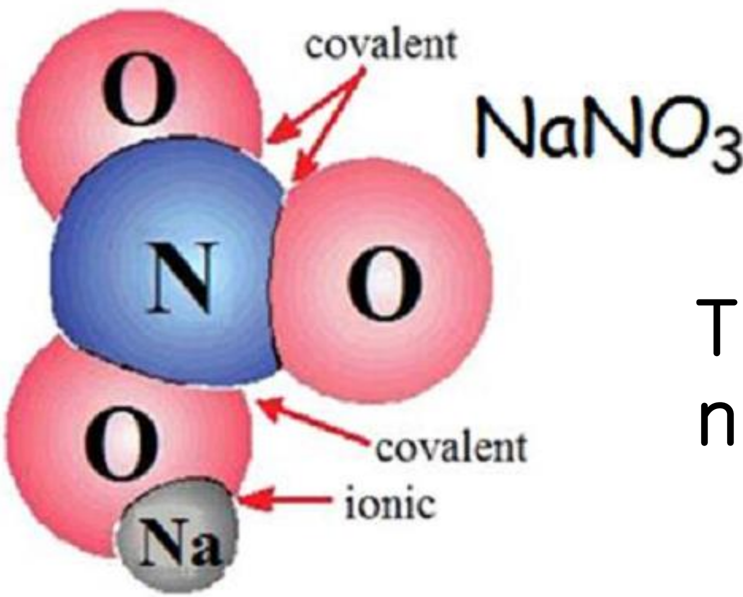


- The molecule of glucose can be reacted with six oxygen atoms to produce six molecules of carbon dioxide and six molecules of water.
- During the reaction, the atoms of the glucose molecule are rearranged into the structure of carbon dioxide and water molecules.
- The bonds in the glucose are broken and new bonds are formed.

- As this occurs, potential energy is released because the new bonds have lower potential energy that the original bonds.
- The bonds in the products are lower energy bonds and therefore, the product molecules are more stable.

Some Compounds Have Both Covalent and Ionic Bonds

- Once the polyatomic ion is constructed with covalent bonds, it reacts with other substances as an ion.
- The bond between a polyatomic ion and another ion will be ionic.



The bonding in sodium nitrate, NaNO3.

Summary

- Covalent bonds are formed by electrons being shared between two atoms.
- Half-filled orbitals of two atoms are overlapped and the valence electrons shared by the atoms.
- Bond energy is the amount of energy necessary to break the covalent bond.
- The strength of a covalent bond is measured by the bond energy.
- Stable compounds have high bond energy and unstable compounds have low bond energy.

Review Questions

1. Describe the characteristics of two atoms that would be expected to form an ionic bond.
2. Describe the characteristics of two atoms that would be expected to form a covalent bond.
3. If an atom had a very high bond energy, would you expect it to be stable or unstable?

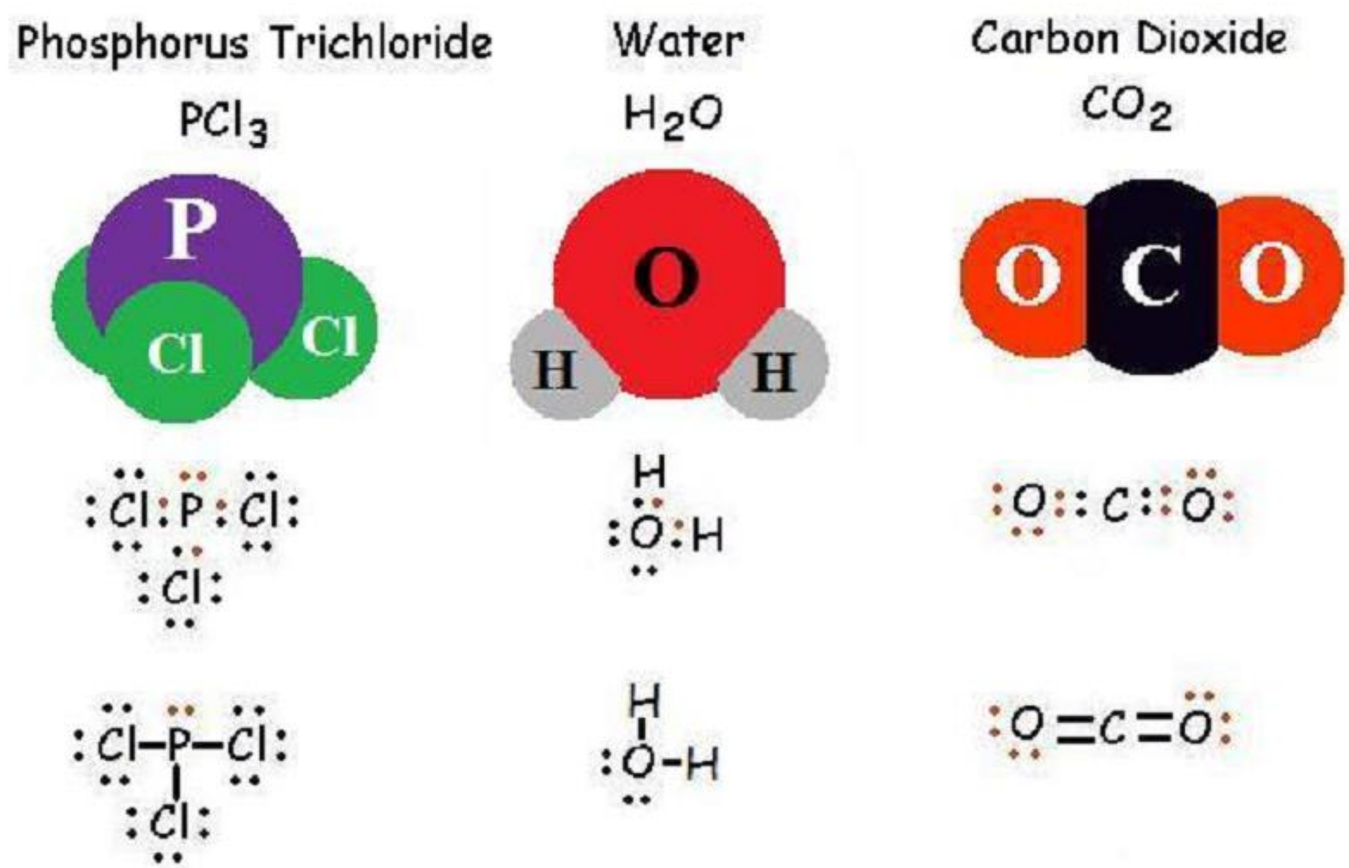
Atoms that Form Covalent Bonds

Introduction

- The bonding of atoms is directly by the laws of nature relating to the tendency toward minimum potential energy, electrical attraction and repulsion, and the arrangement of electrons in atoms.
- As it happens, these laws of nature and energy conditions do favor (in most cases) an octet of electrons for atoms.
- In ionic bonding, the atoms acquired this octet by gaining or losing electrons and, in covalent bonding, as you have seen, the atoms acquire the noble gas electron configuration by sharing electrons..

Non-Metals Bond with Non-Metals to Form Covalent Bonds

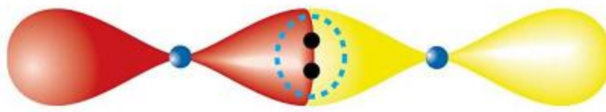
- Metals simply do not hold on to electrons with enough strength to form much in the way of covalent bonds.
- For a covalent bond to form, we need two atoms that both attract electrons with high electron affinity.
- Hence, the great majority of covalent bonds will be formed between two non-metals.



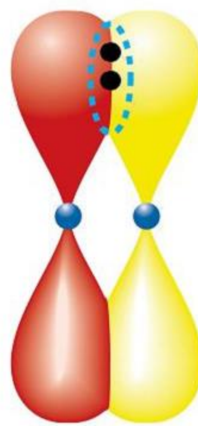
Various methods of showing a covalent bond.

Multiple Bonds

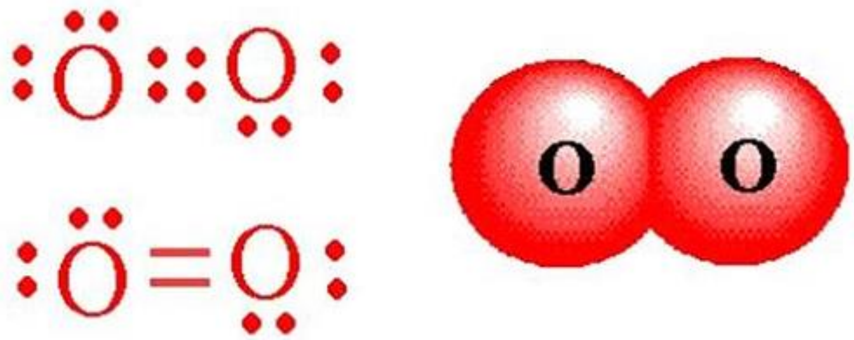
- Covalent bonds in which one pair of electrons is shared; This type of bond is called a single bond.
- When atoms share two pairs of electrons, it is called a double bond and when atoms share three pairs of electrons, it is called a triple bond.



The end-to-end overlap of orbitals forms a sigma bond.



The side-to-side overlap of orbitals forms a pi bond.



Oxygen is double bonded.

Nitrogen is triple bonded

Double bonds

- Double bonds are stronger than single bonds.
- They are not exactly twice as strong; sometimes they are more than twice as strong and sometimes they are less than twice as strong.
- Oxygen is a reactive element and it is surprising that there is so much elemental oxygen in our atmosphere.
-

Double bonds

- The explanation for the existence of so much elemental oxygen is that the double bond is very strong and it takes a great deal of energy to break the double bond in oxygen so that the oxygen atoms could react with something else.

Triple bonds

- The nitrogen molecule, N2, is triple bonded.
- That means that the two nitrogen atoms share three pairs of electrons.
- The first bond will be an end-to-end overlap (**σ** bond) and the other two bonds will be side-to-side overlaps (**π** bonds).
- If the end-to-end overlap are the px orbitals, then the side-to-side overlaps will be the py and pz orbitals.

Lewis Formulas

- What was called “electron dot formulas” when drawing themfor individual atoms become “Lewis dot formulas” or “Lewis structures” or “Lewis formulas” when drawing them for molecules.
- The Lewis structures of a molecule show how the valence electrons are arranged among the atoms of the molecule.

Rules for Writing Lewis Structures

- Decide which atoms are bonded.
- Count all the valence electrons of all the atoms.
- Place two electrons between each pair of bonded atoms.
- Complete all the octets (or duets) of the atoms attached to the central atom.
- Place any remaining electrons on the central atom.
- If the central atom does not have an octet, look for places to form double or triple bonds.

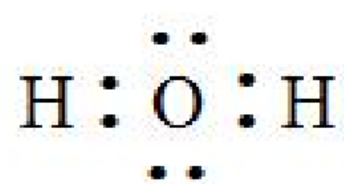
Example 1:

- Write the Lewis structure for water, H_2O.
- Step 1: Decide which atoms are bonded.
- Begin by assuming the hydrogen atoms are bonded to the oxygen atom. i.e. Assume the oxygen atom is the central atom. H – O – H.
- Step 2: Count all the valence electrons of all the atoms.
- The oxygen atom has 6 valence electrons and each hydrogen has 1. The total is 8.
- Step 3: Place two electrons between each pair of bonded atoms.

H : O : H

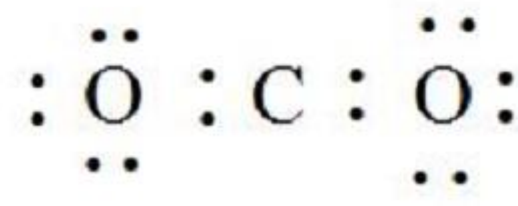
Example 1:

- <u>Step 4</u>: Complete all the octets or duets of the atoms attached to the central atom.
- The hydrogen atoms are attached to the central atom and hydrogen atoms require a duet of electrons and those duets are already present.
- <u>Step 5</u>: Place any remaining electrons on the central atom.



Example 2:

Write the Lewis structure for carbon dioxide, CO2.



Is this structure correct?

Is the total number of valence electrons correct?

Does each atom have the appropriate duet or octet of electrons?

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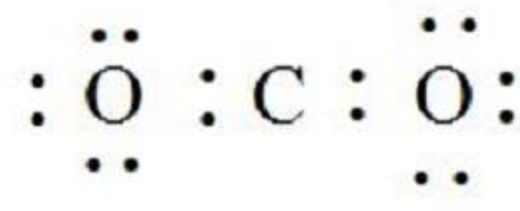
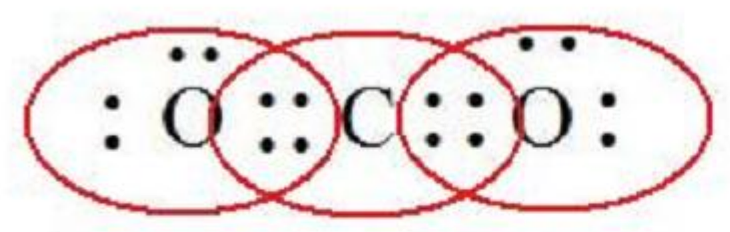


Is this structure correct?

Is the total number of valence electrons correct? Yes

Does each atom have the appropriate duet or octet of electrons? NOT correct

- Step 6: If the central atom does not have an octet, look for places to form double or triple bonds.
- Double bonds can be formed between carbon and each oxygen atom.

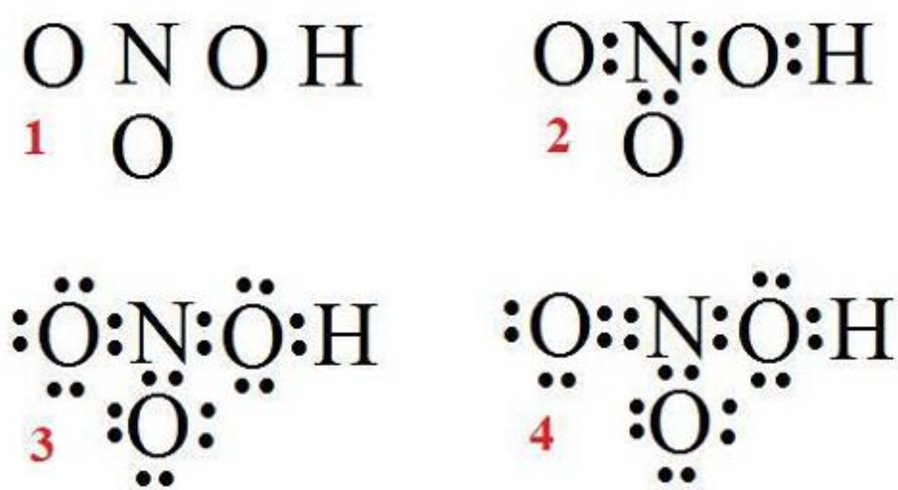


Example 3:

Write the Lewis structure for ammonia, NH3.

Example 4:

Given the skeleton structure for nitric acid, HNO3, place the electrons into a proper Lewis structure.



Summary

- Covalent bonds are formed between atoms with relatively high electron affinity.
- Some atoms are capable of forming double or triple bonds.
- Multiple bonds between atoms require multiple half-filled orbitals.

Summary

- End-to-end orbital overlaps are called sigma bonds.
- Side-to-side orbital overlaps are called pi bonds.
- Lewis structures are commonly used to show the valence electron arrangement in covalently bonded molecules.

Review Questions

1. Which of the following compounds would you expect to be ionically bonded and which covalently bonded?

Compound	Ionic or Covalent
CS_2	
K_2S	
FeF_3	
PF_3	
BF_3	
AlF_3	
BaS	

2. How many sigma bonds and how many pi bonds are present in a triple bond?

3. Draw the Lewis structure for CCl

4. Draw the Lewis structure for SO2.

Naming Covalent Compounds

The Number of Atoms in the Formulas Must be Indicated

- In naming ionic compounds, there is no need to indicate the number of atoms of each element in a formula because, for most cases, there is only one possible compound that can form from the ions present.
- When aluminum combined with sulfur, the only possible compound is aluminum sulfide, Al2S3.

- The only exception to this is a few variable oxidation number metals and those are handled with Roman numerals for the oxidation number of the metal, as in iron (II) chloride, FeCl2.

- With covalent compounds, however, we have a very different situation.
- There are six different covalent compounds that can form between nitrogen and oxygen and in two of them, nitrogen has the same oxidation number.
- Therefore, the Roman numeral system will not work.
- Chemists devised a nomenclature system for covalent compounds that indicate how many atoms of each element is present in a molecule of the compound.

Greek Prefixes

Greek Pre-
fixes

Prefix	**Number Indicated**
Mono-	1
Di-	2
Tri-	3
Tetra-	4
Penta-	5
Hexa-	6
Hepta-	7
Octa-	8
Nona-	9
Deca-	10

Greek Prefixes

Examples

N_2O	dinitrogen monoxide
NO	nitrogen monoxide
NO_2	nitrogen dioxide
N_2O_3	dinitrogen trioxide
N_2O_4	dinitrogen tetraoxide
N_2O_5	dinitrogen pentaoxide
SF_6	sulfur hexafluoride
CO_2	carbon dioxide
P_4O_{10}	tetraphosphorus decaoxide
P_2S_5	diphosphorus pentasulfide

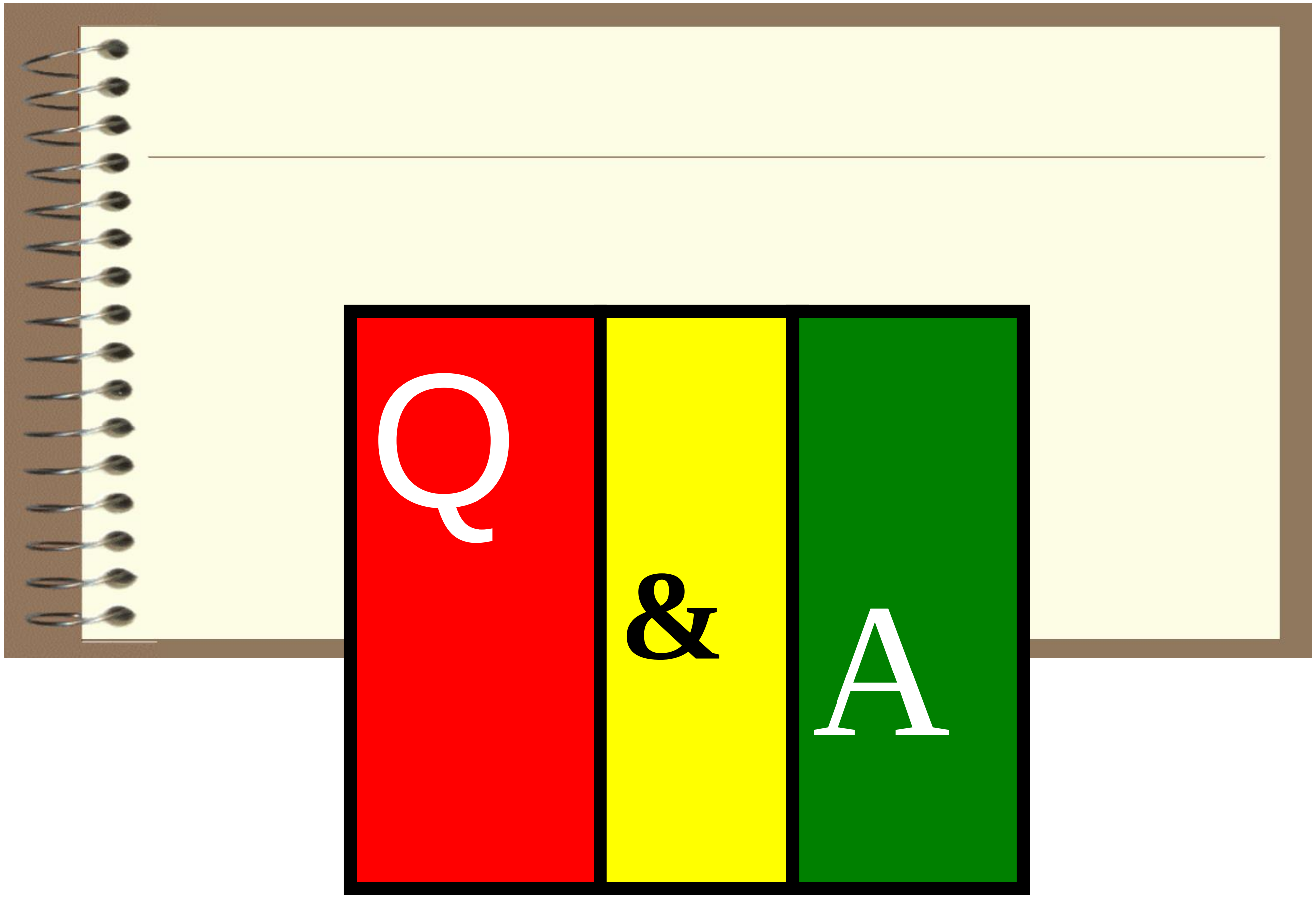
Questions

1. Name the compound CO.
2. Name the compound PCl

3. Name the compound PCl

4. Name the compound N2O3.
5. Name the compound BCl

6. Name the compound SF4.
7. Name the compound Cl2O.



Thanks.