Chemical bonding & structure

Ionic bonding and structure

Covalent bonding

Covalent structures

Intermolecular forces

Metallic bonding

Ms. Thompson - SL Chemistry
Wooster High School

Topic 4.2
Covalent bonding

A covalent bond is formed by the electrostatic attraction between a shared pair of electrons and the positively charged nuclei.

- Single, double and triple covalent bonds involve one, two and three shared pairs of electrons respectively.
- Bond length decreases and bond strength increases as the number of shared electrons increases.
- Bond polarity results from the difference in electronegativities of the bonded atoms.

Covalent bonding

Nature of science

- Looking for trends and discrepancies – compounds that contain nonmetals have different properties from compounds that contain nonmetals and metals
- Use theories to explain natural phenomena – Lewis introduced a class of compounds which share electrons. Pauling used the idea of electronegativity to explain unequal sharing of electrons.

Covalent bonding

Covalent bonding

- A **covalent bond** is formed by the electrostatic attraction between a shared pair of electrons and the positively charged nuclei.
  - Atoms *share electrons* with each other in order to attain noble gas configuration. Usually occurs between *nonmetals*.
- In order to understand the covalent bonding it is important to first understand how to do a **Lewis (electron dot) structure**. This is a convenient method of representing the valence (outer shell) electrons of an element.
  - Each element is surrounded by a number of dots (or crosses) to represent the valence electrons of the element.

Covalent bonding

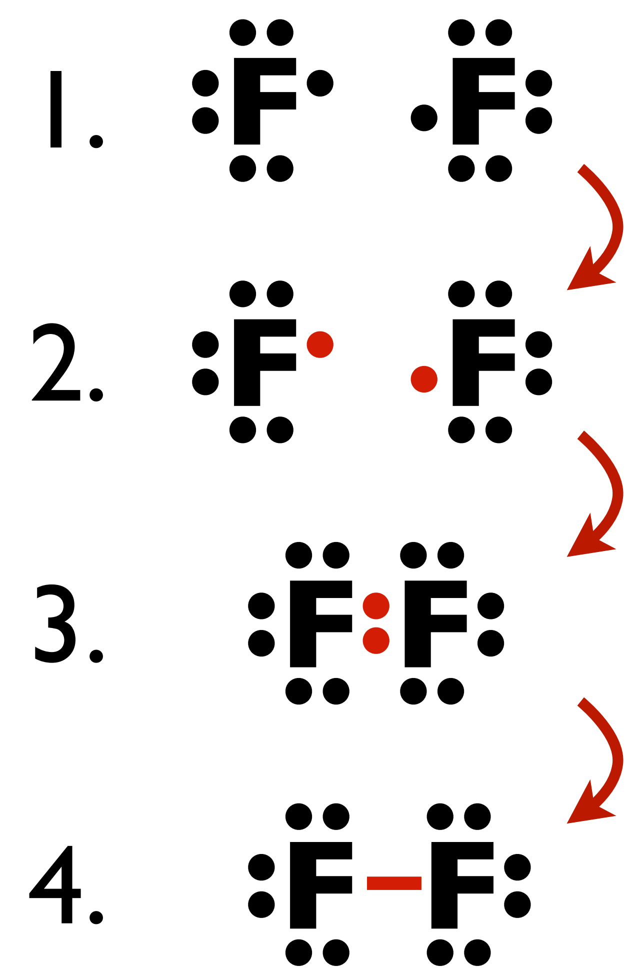
Lewis dot structures

1. Count the number of valence electrons for each atom in the molecule
   a) Group number
   b) Adjust valence electrons if drawing structure of an ion
2. Draw atoms bonded together
   a) **Central** atom is drawn first in molecule
   b) Each bond contains **two** electrons
3. Fill in remaining valence electrons to give all atoms an octet of electrons leaving central atom last
   a) If not enough electrons use multiple bonds
   b) If excess electrons, place around central atom

Covalent bonding

Lewis Dot Structure

- **Fluorine, F_2**
  - Group 17 and has seven valence electrons
  - Needs to acquire one more electron for a full octet and to attain noble gas configuration.
    - If two fluorine atoms share one electron with each other, each fluorine atom gains one more electron to attain a complete octet of electrons, which results in the formation of a covalent bond between the two fluorine atoms.
    - A **single, covalent bond** is formed and the shared pair of electrons can be represented by a line.



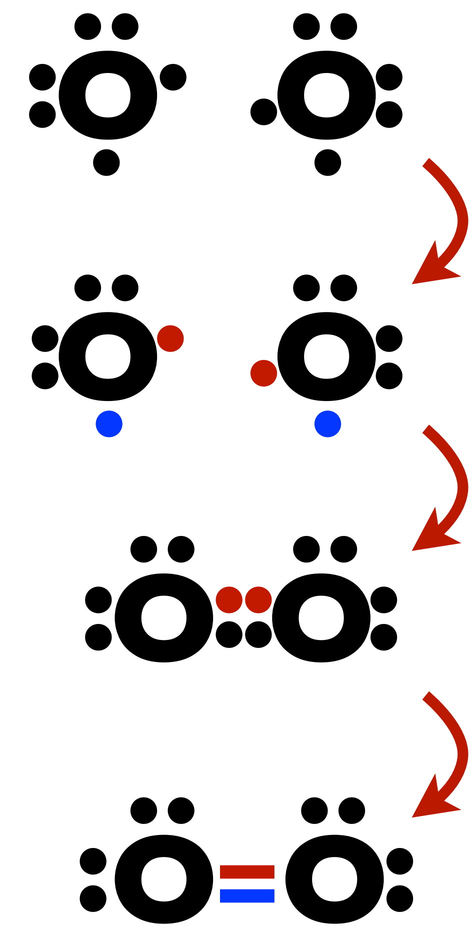
Fluorine has 6 non-bonding pair of electrons (lone pairs) and one pair that bonds

Covalent bonding

Lewis Dot Structure

- **Oxygen, O_2**
  - Group 16 and has six valence electrons
  - Needs to acquire two more electrons for a full octet and to attain noble gas configuration.
    - If two oxygen atoms share two electrons with each other, each oxygen atom gains two more electrons to attain a complete octet of electrons, which results in the formation of a covalent bond between the two oxygen atoms.
    - A **double, covalent bond** is formed and the shared pair of electrons can be represented by two lines.

1. 2. 3. 4.

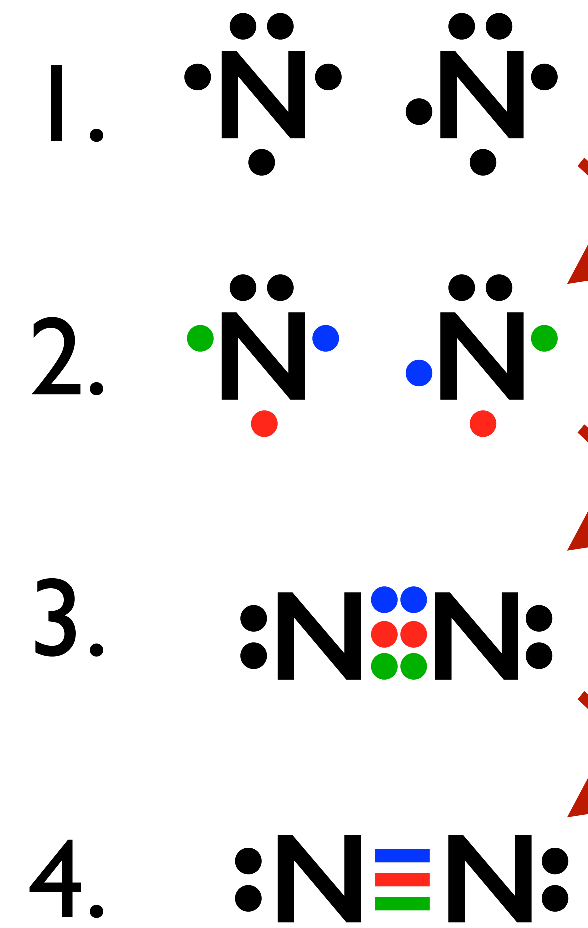


Oxygen has 4 non-bonding pair of electrons (lone pairs) and two pair that bonds

Covalent bonding

Lewis Dot Structure

- **Nitrogen, N_2**
  - Group 15 and has five valence electrons
  - Needs to acquire three more electrons for a full octet and to attain noble gas configuration.
    - If two nitrogen atoms share three electrons with each other, each nitrogen atom gains three more electrons to attain a complete octet of electrons, which results in the formation of a covalent bond between the two nitrogen atoms.
    - A **triple, covalent bond** is formed and the shared pair of electrons can be represented by three lines.

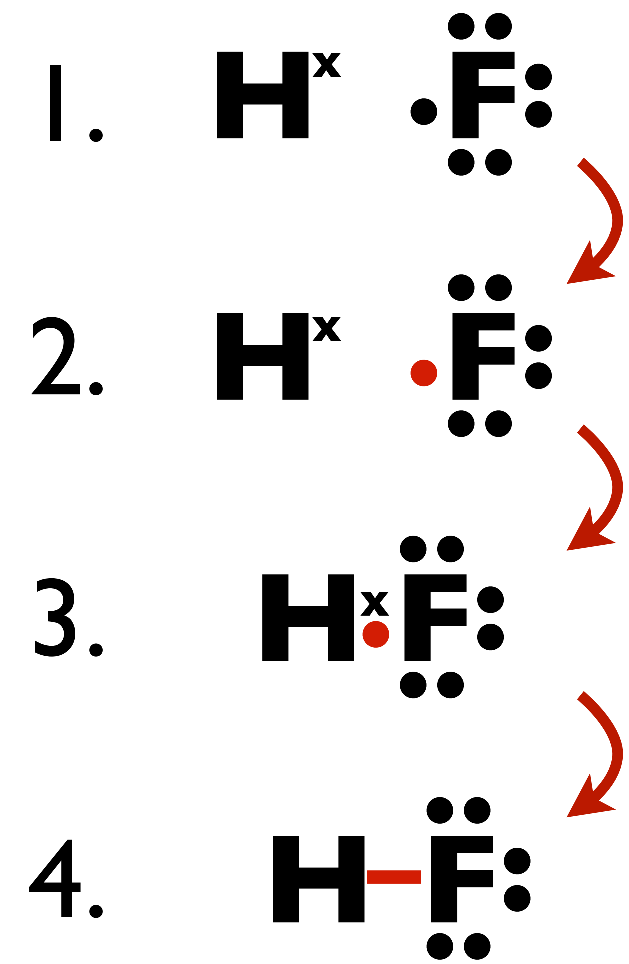


Oxygen has 4 non-bonding pair of electrons (lone pairs) and two pair that bonds

Covalent bonding

Lewis Dot Structure

- **Hydrogen Fluoride, HF**
  - Fluorine is in group 17 and has seven valence electrons and hydrogen has one
  - Both need to acquire one more electron to attain noble gas configuration (fluorine will gain a full octet).
    - If hydrogen and fluorine atoms share one electron with each other, both atoms gain one more electron to attain a full outer shell of electrons (H, 2 and F, 8), which results in the formation of a covalent bond between the two atoms.
    - A **single, covalent bond** is formed and the shared pair of electrons can be represented by a line.



Fluorine has 3 non-bonding pair of electrons (lone pairs) and one pair that bonds

Practice Problem

20 mins

... You Do ...

Work with a partner and complete the lewis dot structure handout

Lewis structures

Review of steps to drawing Lewis structures:

1) Calculate number of valance electrons in the molecule.
2) Calculate the number of electrons each atom needs to complete its octet.
3) Subtract 1 from 2 – this will give you the number of bonding electrons in the molecule.
4) Draw the skeletal structure of the molecule with the least electronegative atom at the center.
5) Complete the octets of the atoms in the molecule.
6) Add multiple bonds to complete the octets (if necessary).

Covalent bonding

Bond strength and bond length

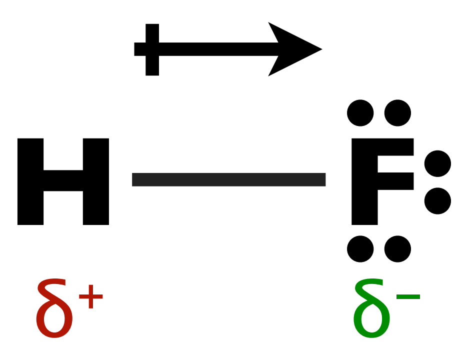
- Single, double, and triple bonds differ in both strength and length
- **Bond strength**
  - Triple > double > single
  - *Found in section 11 of data booklet*
- **Bond length**
  - single > double > triple
  - *Found in section 10 of data booklet*

Bond	**Bond enthalpy (at 298K) kJ mol^{-1}**	**Covalent bond length / pm**
C≡C	839	120
C=C	614	134
C−C	346	154

Covalent bonding

Comparison of covalent bonds and ionic bonds: Electronegativity

- Some atoms have a stronger pull for electrons (electronegativity)
- In identical atoms, the sharing of electrons is equal but in other cases (such as HF) there is an unequal sharing of electrons. This creates a **polar covalent bond**.
- One atom will adapt a partial negative charge (δ^-) and the other will adapt a partial positive charge (δ^+)
  - *In HF, fluorine is more electronegative than hydrogen so it pulls more electrons towards it. So fluorine acquires a partial negative charge, δ^- and hydrogen acquires a partial positive charge, δ^+.*
- This separation of charge can be represented vectorially **dipole moment, μ**.



Covalent bonding

Comparison of covalent bonds and ionic bonds: Electronegativity

- If atoms are identical in the formation of a covalent bond, it is said to be a **pure covalent bond**.
  - i.e. F_2
  - There is no dipole moment so this is a **non-polar covalent bond**
  - **Electronegativity** is the relative attraction that an atom of an element has for a shared pair of electrons.
  - Electronegativity values in section 8 of data booklet

Trends in electronegativities

•Going from L to R across a period electronegativity values *increase*.

Reason:

- •decreasing atomic radii
- •increasing nuclear charge

•Going down a group electronegativity values *decrease*.

•*Reason*:

- •increasing atomic radii
- •shielding effect of inner electrons

KNOW

Covalent bonding

Electronegativity rules

- Benefit of knowing the electronegativity values, χ_p, is that we can estimate, based on electronegativity value differences, $\Delta\chi_P$, whether a bond is ionic, pure covalent (non-polar), or polar covalent.
- Based on the following rules you should **know**:

Bond type	$\mathbf{\Delta\chi_P}$
ionic	$\Delta\chi_P > 1.8$
pure covalent (non-polar)	$\Delta\chi_P = 0$
polar covalent	$0 < \Delta\chi_P \leq 1.8$

Differences between ionic and covalent bonding

Ionic bonding	Covalent bonding
Formed between a cation (usually metal) and an anion (usually nonmetal). Some cations (such as NH_4^+) can be comprised of non-metals and anions (such as MnO_4–) can contain metals.	Usually formed between nonmetals
Forms from atoms either losing electrons (process of oxidation) or gaining electrons (process of reduction) in order to obtain a noble gas electron configuration.	Formed from atoms sharing electrons with each other in order to attain a noble gas electron configuration.
Electrostatic attraction between oppositely charged ions, that is, a cation (positive ion) and an anion (negative ion).	Electrostatic attraction between a shared pair of electrons and the positively charged nuclei.
Ionic compounds have lattice structures	Covalent compound consist of molecules*
Ionic compounds have higher melting points and boiling points	Covalent compounds have a lower melting points and boiling points
Ionic compounds have low volatilities	Covalent compounds maybe volatile
Ionic compounds tend to be soluble in water	Covalent compounds typically are insoluble in water
Ionic compounds conduct electricity because ions are free to move in the molten state. They do not conduct electricity when solid however, as the ions are not free to move.	Covalent compounds do not conduct electricity because no ions are present to carry the charge.

Practice Problem

... I Do ...

Deduce which of the following compounds are molecular:

a) SO_2
b) PCl_3
c) Na_2O
d) NH_4NO_3

SO_2 and PCl_3 are <u>molecules</u> as they only contain nonmetals and no ions.

Na_2O is <u>ionic</u> because $\Delta\chi_P > 1.8$

NH_4NO_3 is <u>ionic</u> even though it contains all non-metals. It contains the ammonium ion NH_4^+ and the nitrate oxoanion (ions that contain oxygen), NO_3^-.

** Remember that ionic compounds have lattice structures*

Practice Problem

... You Do ...

Deduce which of the bonds in the following binary compounds are ionic, pure covalent (non-polar), or polar covalent:

a) H_2
b) HCl
c) KBr
d) CO

Practice Problem

... You Do ...

Deduce which of the bonds in the following binary compounds are ionic, pure covalent (non-polar), or polar covalent:

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b) HCl
c) KBr
d) CO

H_2 $\Delta\chi_P = 0$, so H_2 is a pure covalent (non-polar) bond

HCl $\Delta\chi_P(Cl) = 3.2$ and $\Delta\chi_P(H) = 2.2$, so $\Delta\chi_P = 1.0$, so HCl has a polar covalent bond, with chlorine having partial negative charge, δ^-, and hydrogen having a partial positive charge, δ^+.

KBr $\Delta\chi_P(Br) = 3.0$ and $\Delta\chi_P(K) = 0.8$, so $\Delta\chi_P = 2.2$, so KBr has ionic bonding with potassium having a 1+ charge and bromine having a 1- charge.

CO $\Delta\chi_P(O) = 3.4$ and $\Delta\chi_P(C) = 2.6$, so $\Delta\chi_P = 0.8$, so CO has a polar covalent bond, with oxygen having a partial negative charge, δ^-, and carbon having a partial positive charge, δ^+.

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Covalent bonding

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