

McGill **CHEM 110**

Fall 2024, Chapter 3 Notes



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3. Models of Chemical Bonding

3.1 Types of Bonds

3.1.1

Introduction to Intramolecular Bonds

Your university just had a poster sale! You bought a new poster to add to your collection and you were trying to decide what to use to stick the poster to the wall. You could have used:



Just like how we can use different things to stick the poster to the wall, there are different bonds that connect different atoms together.

Intramolecular bond: a bond that connects two atoms within a molecule together

Before we take a look at the different types of intramolecular bonds, we will review electronegativity.

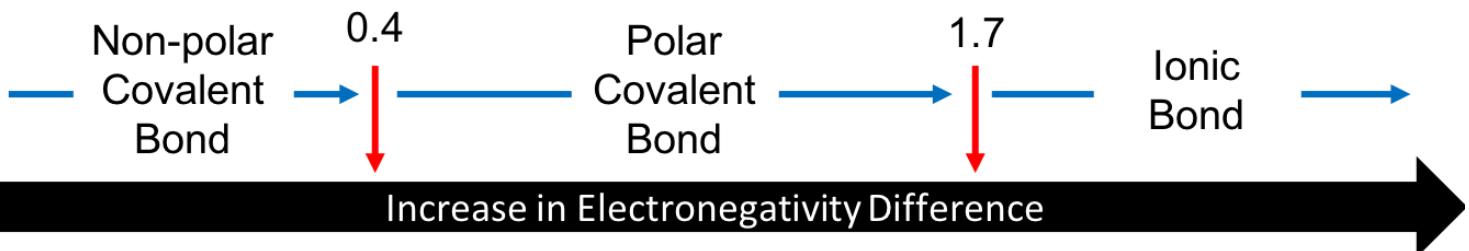
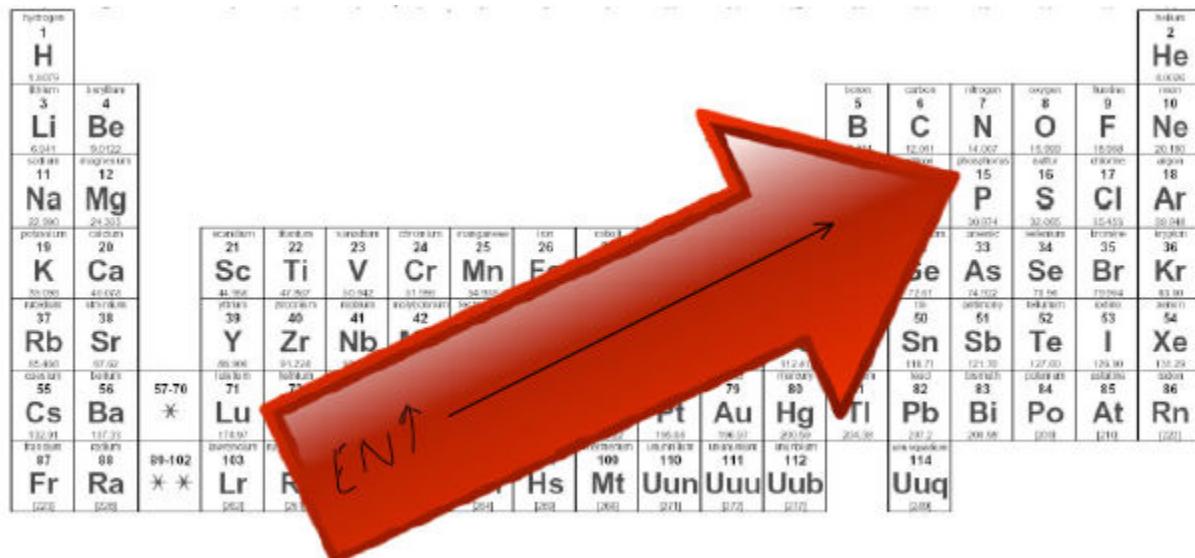
Electronegativity: is the tendency of an atom to pull bonding electrons towards itself

- Valence electrons are involved in chemical bonding
- The type of bond depends on the difference in electronegativity between the two atoms in the bond

! **WATCH OUT!**

This is similar to electron affinity but not the same!

Electron affinity involves a single atom/ion, whereas **electronegativity involves two bonded atoms.**



Types of Intramolecular Bonds

We will be discussing ionic bonds, covalent bonds (polar covalent bonds, non-polar covalent bonds, and coordination covalent bonds), as well as metallic bonds, which are tested on the least.

Ionic Bonds:

- Are between a **metal and a non-metal**
- You might see ionic compounds called **salts** like NaCl
- There is a **large difference in EN** ($\Delta\text{EN} > 1.7$) in these bonds
- The **metal gives electrons to the non-metal** (no sharing of electrons)

Example:

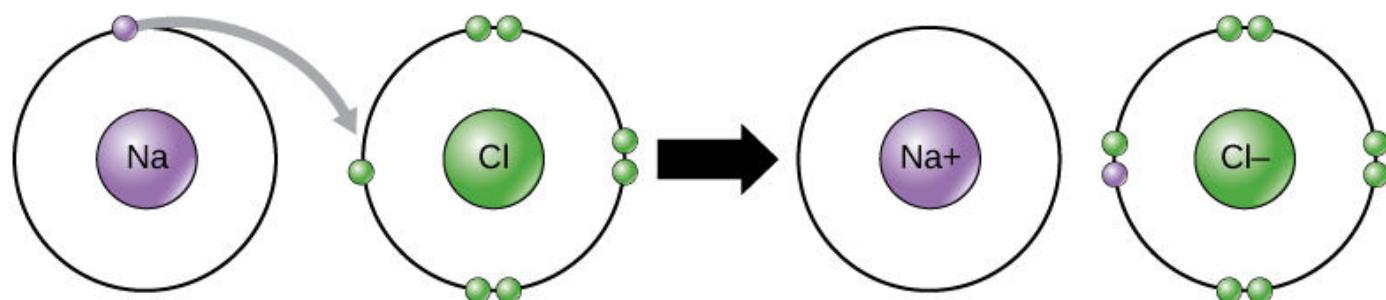


Photo by CNS Openstax/ CC BY

The **charges** that result are what allow ionic compounds to **conduct electric current**

Covalent Bonds:

- Bonds where electrons are **shared** between **two non-metals**
- We will consider 3 types of covalent bonds: i) Non-polar covalent bonds, ii) Polar covalent bonds
iii) Coordination covalent bonds

Non-polar Covalent Bonds:

- Electrons are **shared equally** between **2 of the same non-metals**
- As a result there is a very small difference in EN, $0 < \Delta\text{EN} < 0.4$

Examples: H_2 , O_2 , N_2 , Cl_2 , * S_8 , * P_4

Example: Shown below is an example of two hydrogen atoms sharing their single electron to form a covalent bond. This is specifically a non-polar covalent bond!

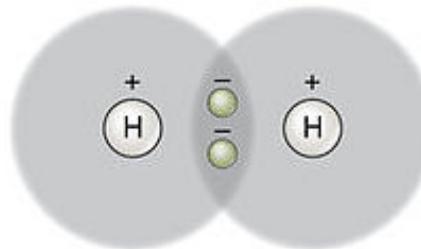


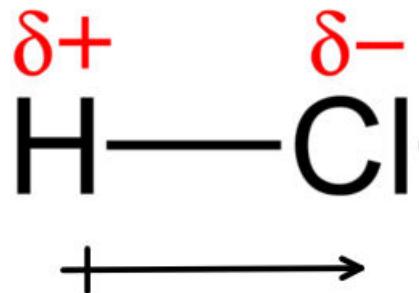
Photo by OpenStax College / CC BY

Polar Covalent Bonds:

- Electrons are shared **unequally** between **2 different non-metals**
- There is a difference in EN, $0.4 < \Delta\text{EN} < 1.7$

Since electrons are shared unequally in this bond, we say that there is a "**dipole moment**"

- The dipole moment is a vector with both magnitude and direction (more on this later!)
 - Partial negative charge (δ^-) is assigned to the atom with the **higher EN**
 - Partial positive charge (δ^+) is assigned to the atom with the **lower EN**
- **The greater the difference in electronegativity (EN), the greater the dipole moment!**



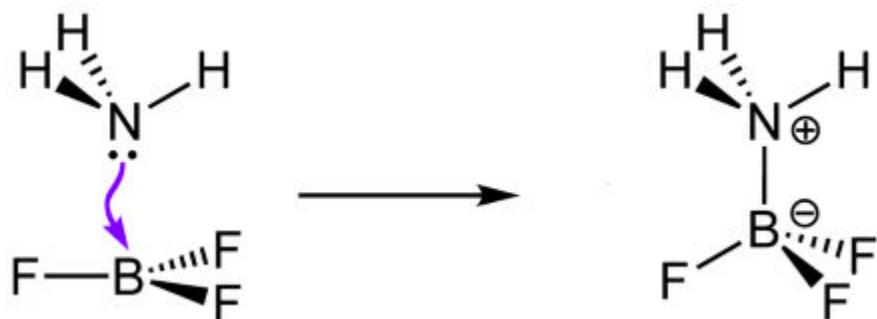
i WIZE TIP

You can think of the dipole moment as a tug of war. The dipole moment (arrow) points towards the winner that is able to pull electrons more towards itself (is more electronegative!)

Coordination Covalent Bonds:

- These are covalent bonds (between 2 non-metals) where **both electrons in the bond are donated by one of the non-metals**

Example: NH_3 reacting with BF_3



- The charges we see here as a result of the bond forming help the product **weakly conduct in solution**
 - **Note:** Ionic bonds conduct a much stronger electric current!

Metallic bond

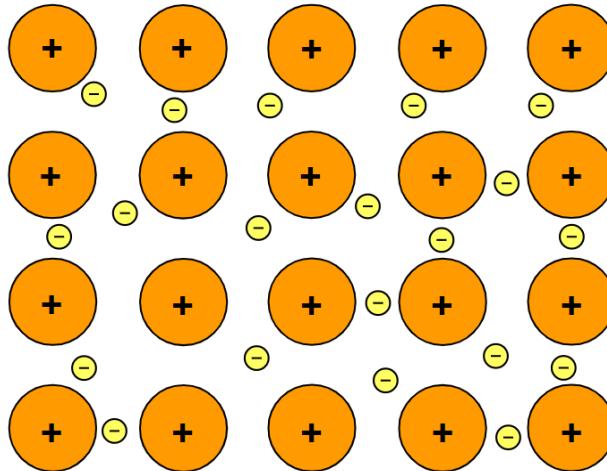


Photo by Muskid / CC BY

- The diagram is showing many atoms of a metal element and their inner shell electrons are surrounded by a **sea of electrons** that are free to move around (aka are delocalized)
 - The **circles with a "+" inside of them** represent a **metal atom + inner shell electrons** (recall nuclei are positively charged)
 - The **smaller circles with a "-" inside of them** represent **valence electrons**
- At least 1 VE/atom is free=**conduction electrons**
- These conduction electrons are what give the metal their properties!
 - Ductile, malleable, conduct thermal energy, conduct electricity, have lustre/shine

Examples: Aluminum metal, Iron, Zinc etc.

3.1.3

Example: Ionic vs Covalent Bonding

For the following list of compounds, determine which are mostly ionic and which are mostly covalent.



Example: Types of Intramolecular Bonds

Indicate what type of bond will be formed between the two atoms in the following compounds. Try to be specific. :)

a) CH_4

b) NH_3

c) LiBr

d) Br_2

e) Na_2SO_4

3.1.5

What type of bonding would be expected in the following compounds?

Part 1

Table salt (NaCl)

Ionic bonding



Polar covalent bond



Non-polar covalent bond



Metallic bond



What type of bonding would be expected in the following compounds?

Part 2

Ammonia (NH_3)

Ionic bonding



Polar covalent bond



Non-polar covalent bond



Metallic bond



What type of bonding would be expected in the following compounds?

Part 3

Nitrogen (N_2)

Ionic bonding

Polar covalent bond

Non-polar covalent bond

Metallic bond

Practice: Types of Intramolecular Bonds

For the three compounds: H_2 , CCl_4 and MgF_2 , indicate which types of bonds each molecule has.

- A.** ionic bond(s)
- B.** polar covalent bond(s)
- C.** non-polar covalent bond(s)

H_2

CCl_4

MgF_2

3.2

Lattice Energy of Ionic Bonds

3.2.1

Lattice Energy

Lattice energy: is the **energy required to break ionic bonds into cation anion pairs**, or the amount of energy released by forming the ionic compound from the constituent ions.



The lattice energy for an ionic compound depends on the strength of the electrostatic interactions between the cation and anion which can be calculated using the equation below.

$$E_{lattice} = k \frac{Q_1 Q_2}{d}$$

E_{lattice} is the **lattice energy**

k is a constant $8.99 \times 10^9 \text{ J m/C}^2$

Q₁ is the **charge of ion 1**

Q₂ is the **charge of ion 2**

d is the **bond length**

i WIZE TIP

On an exam, it is very unlikely that you would see a question that asks you to calculate the lattice energy.

A more typical exam question will ask you to compare the lattice energies between two different ionic compounds and **determine which of the ionic compounds has the higher lattice energy**.

If an ionic bond has a higher lattice energy, it means that ionic bond is (stronger/weaker) _____

$$E_{lattice} = k \frac{Q_1 Q_2}{d}$$

According to the equation for lattice energy, answer the following questions.

1. If the **ions** that are forming the ionic bond have **greater charges**, will the lattice energy be higher or lower? _____
2. If the atomic radii of the atoms forming the ionic bond is larger, the **bond length will be larger** and the lattice energy will be larger or smaller? _____

 **WIZE TIP**

Charge is more important than size when ranking compounds based on their lattice energy.

3.2.2

Example: Lattice Energy

Which of the following pairs of ionic compounds will have stronger ionic bonding (*i.e. higher lattice energy*)

Group →	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
↓ Period																		
1	1 H																2 He	
2	3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
3	11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
4	19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
5	37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
6	55 Cs	56 Ba		72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
7	87 Fr	88 Ra		104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 Cn	113 Nh	114 Fl	115 Ms	116 Lv	117 Ts	118 Og
Lanthanides		57 La	58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu		
Actinides		89 Ac	90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr		

a) LiF vs BeO

b) LiF vs CsF



c) LiF vs MgS

3.2.3

Example: Ranking Lattice Energy

Rank the following compounds in order of increasing lattice energy, RbCl, CaI₂, BaSe.

Practice: Ionic Bond Strength

Consider two compounds: CaS and NaCl. Which ionic compound would have a larger Lattice Energy released upon forming their respective ionic compounds. Why?

CaS, because the two ions are larger than the ions in LiCl.

CaS, because the two ions have double the charge compared to the ions in LiCl.

CaS, because the two ions are smaller than the ions in LiCl.

LiCl, because the two ions are larger than the ions in CaS.

LiCl, because the two ions have smaller charges than the two ions in CaS.

LiCl, because the two ions are smaller than the ions in CaS.

3.2.5

Practice: Highest Lattice Energy

Choose the option below with the highest lattice energy.

LiCl

CO₂

MgO

CaS

MgI₂

3.3

Properties of Covalent Bonds

3.3.1

Properties of Covalent Bonds

Bonds can have various bond lengths and bond energies. There can also be a different number of bonds between different atoms. We will take a look at some of these terms and properties of bonds now.

Bond Length

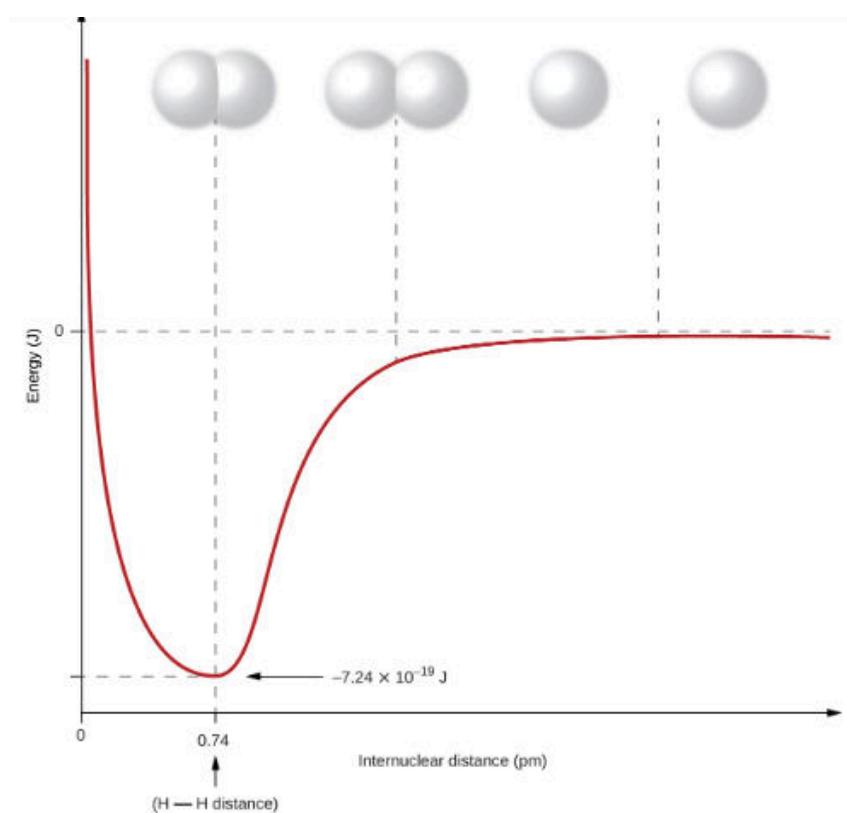


Photo by Rice University / CC BY

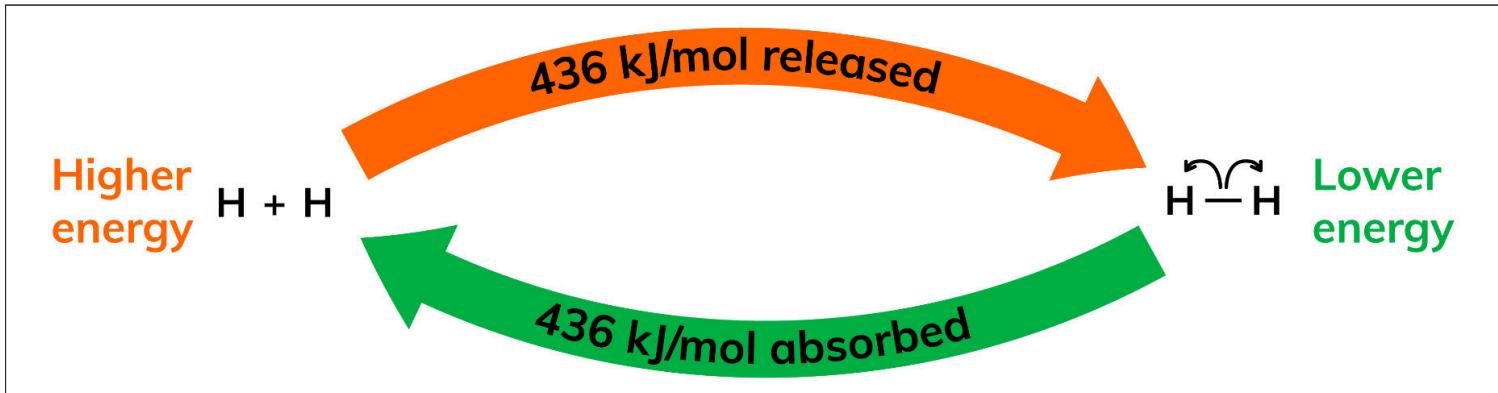
- When 2 atoms get bound together, there are repulsion forces and attractive forces
 - The electrons from one atom are attracted to the other atom's nucleus (attraction)
 - But the nuclei from both atoms can't get too close (repulsion because of + charges)
 - When the forces balance out we form a bond!

Circle the most stable bond in the diagram above (this is how the bond will exist).

Bond length: this is the **distance between two nuclei** that are bound together when they are at their **lowest possible energy state**

- **Note:** In the graph, energy rises when atoms have a shorter or longer bond length than what is optimal
- When atoms with larger atomic radii bind together, the bond length (distance between their nuclei) is (shorter/longer) _____

Bond Dissociation Energy (BDE)



Bond dissociation energy: is the amount of energy needed to break a bond homolytically

In the diagram, this is represented by the blue arrow. Indicate where the BDE can be found on the graph above.

WIZE CONCEPT

Breaking bonds will always require energy!

Forming bonds on the other hand will release energy since bonds are stable and lower in energy!

The higher the BDE the stronger/weaker the bond: _____

Bond Order

Bond order: the number of bonds between adjacent atoms

Examples:

The bond order of H-F is _____

The bond order of C=O is _____

The bond order of C≡C is _____

How are Bond Order, Bond Length, and Bond Energy Related?

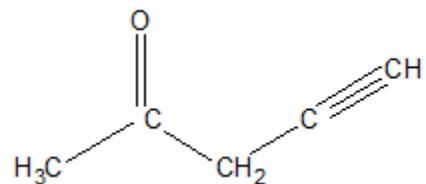
- For bonds between the same elements, the **higher the bond order, the shorter the bond length and the stronger the bond (higher BDE)**
- Bond length is given in units angstroms (\AA) where $1 \text{\AA} = 10^{-10} \text{m}$

Average Bond Lengths and Bond Energies for Some Common Bonds		
Bond	Bond Length (\AA)	Bond Energy (kJ/mol)
C–C	1.54	345
C = C	1.34	611
C ≡ C	1.20	837
C–N	1.43	290
C = N	1.38	615
C ≡ N	1.16	891
C–O	1.43	350
C = O	1.23	741
C ≡ O	1.13	1080

3.3.2

Example: Bond Order

What is the bond order of each bond in this molecule?



Practice: Potential Energy Diagrams

The following figure shows the potential energy diagrams for the carbon-carbon bond in ethane, ethene, and ethyne. Structures of these molecules are shown below. Identify which molecule corresponds to each potential energy curve.

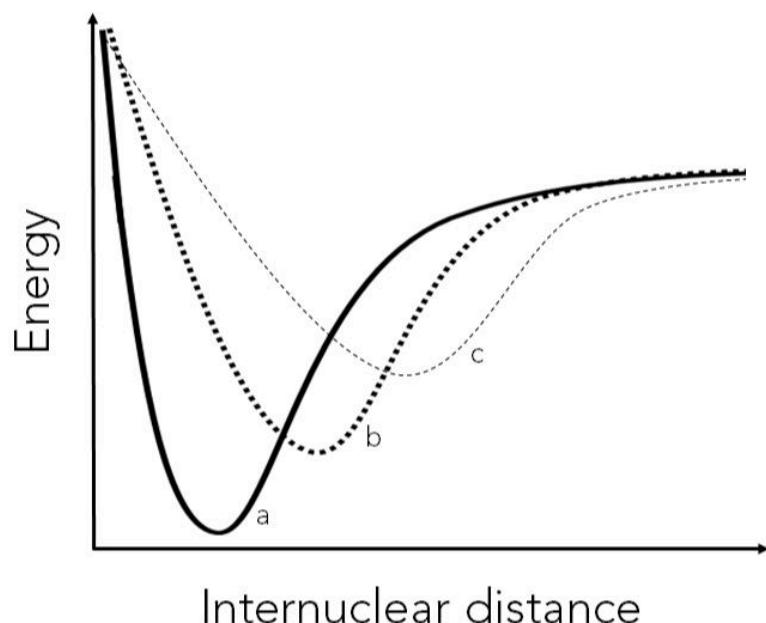
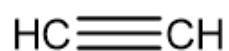
ethane



ethene



ethyne



A. ethene

B. ethyne

C. ethane



potential energy diagram a



potential energy diagram b



potential energy diagram c

3.4 BDE Calculation

3.4.1

Bond Energies

Enthalpy (H) is the energy stored in bonds.

Bond dissociation energy (BDE) is the energy needed to break a chemical bond.

$$\Delta H_{rxn} = [(\sum nH_{bonds\ broken}) - (\sum nH_{bonds\ formed})]$$

Or (an easier way to think about it):

$$\boxed{\Delta H_{rxn} = [(\sum nBDE_{reactants}) - (\sum nBDE_{products})]}$$

ΔH_{rxn} = enthalpy change for the reaction (kJ/mol)

BDE = bond energy per mole of bonds (kJ/mol), always positive

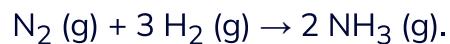
$$\Delta H_{rxn} = [(\sum nBDE_{reactants}) - (\sum nBDE_{products})]$$

Bond Energies (kJ/mol)					
Bond	Bond Energy	Bond	Bond Energy	Bond	Bond Energy
H-H	436	C-S	260	F-Cl	255
H-C	415	C-Cl	330	F-Br	235
H-N	390	C-Br	275	Si-Si	230
H-O	464	C-I	240	Si-P	215
H-F	569	N-N	160	Si-S	225
H-Si	395	N=N	418	Si-Cl	359
H-P	320	N≡N	946	Si-Br	290
H-S	340	N-O	200	Si-I	215
H-Cl	432	N-F	270	P-P	215
H-Br	370	N-P	210	P-S	230
H-I	295	N-Cl	200	P-Cl	330
C-C	345	N-Br	245	P-Br	270
C=C	611	O-O	140	P-I	215
C≡C	837	O=O	498	S-S	215

Photo by Rice University / CC BY

3.4.2

Predict the overall change in enthalpy (i.e. change in enthalpy, change in heat, ΔH_{rxn} in kJ/mol) for the reaction:



BDE:

$$NN = 945 \text{ kJ/mol}$$

$$H-H = 436 \text{ kJ/mol}$$

$$N-H = 391 \text{ kJ/mol}$$

-93

87

1077

1209

-869

3.5

How to Draw Lewis Structures

3.5.1

Drawing Lewis Structures

Drawing the "best" Lewis structure for a molecule is a great skill to have for your chemistry class! This concept will come in handy in the upcoming chapters as well. We will follow the steps below each time.

WIZE TIP

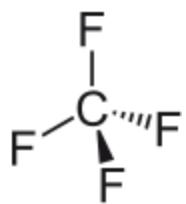
Steps for Drawing Lewis Structures:

- 1) Calculate the total number of **valence electrons** for the molecule.
- 2) Write out all atoms, with the **least electronegative atom in the middle (but H is never in the middle)**
- 3) Connect all atoms with **single bonds**.
- 4) Put **lone pairs** on atoms, except H, until you run out of electrons. Put extra lone pairs on the central atom.
- 5) Shift lone pairs to make double or triple bonds to satisfy the **Octet Rule** and get the best **formal charges**.

Octet Rule: atoms need to have 8 electrons in the valence shell

Examples: C, N, O, F all need 8 electrons in their valence shell to have the "best" Lewis structure

Example: $\text{CF}_4 \rightarrow$ the C atom has 8 valence electrons and each F atom has 8 valence electrons in the Lewis Structure

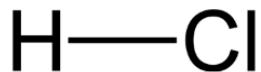


Exceptions to the Octet Rule

Some elements are **octet deficient** (require less than 8 electrons in their valence shell)

Examples:

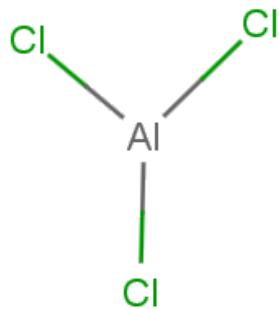
H and He can only have **2 electrons**



Be can only have **4 electrons**



In group 13, Al and B can only have **6 electrons**

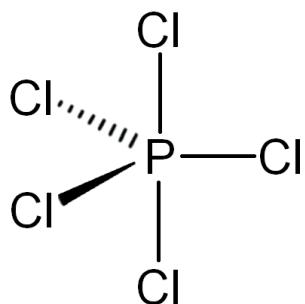


Some elements have an **expanded octet** (can have more than 8 electrons in their valence shell)

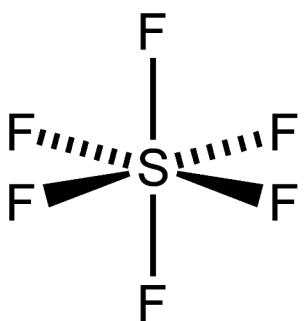
Examples:

Non-metals in period 3 and onwards can have expanded octet.

P can have **10 electrons**



S can have **12 electrons**



Some molecules like **radicals** have an **odd number of electrons** (these are rare!)

Example: NO

First let's count the number of valence electrons (VE) in both atoms.

- N → 5 VEs
- O → 6 VEs
 - So in total there are 11 VEs which is an odd number. We would be left with a molecule that looks like this:



Note: There is an **unpaired electron on the N**. This is called a **radical** molecule and is highly reactive so it is very short-lived.

Formal Charges

The last thing we need to check to ensure we have the **BEST Lewis Structure** is to count formal charges.

WIZE CONCEPT

The best formal charge for an atom is zero.

If the formal charges for a molecule can't be 0 the **best Lewis Structure for that molecule will have the lowest possible formal charges.**

If formal charge is not zero, consider assigning **negative formal charge to more electronegative elements** and positive formal charges to less electronegative elements, if possible.

$$FC = VE - \text{bonds connected to the atom} - \text{lone pair electrons on the atom}$$

FC is the **formal charge** of an atom

VE are the **valence electrons** the atom has

Example:



Example: Lewis Structures

Draw Lewis structures of the following molecules.

a) PCl_3

b) NH_4^+

c) SeF_4

3.5.3

Example: Lewis Structure

Draw the Lewis Structure for SO_4^{2-}

3.5.4

Practice: Formal Charges and Lewis Structures

Draw the Lewis structure of SO_2Cl_2 to answer the following questions.

What is the formal charge on sulfur?

What is the formal charge on each chlorine?

What is the formal charge on each oxygen?

How many double bonds are there in the molecule?

How many single bonds are there in the molecule?

Practice: Drawing Best Lewis Structures

Complete the “best” Lewis diagrams for the following species and use them to answer the next three questions



Part 1

Which Lewis diagram has double bond(s)?

AlI_3

PO_2^+

SeCl_2

all of the above

none of the above

Practice: Drawing Best Lewis Structures

Complete the “best” Lewis diagrams for the following species and use them to answer the next three questions



Part 2

Which Lewis diagram has an atom that exceeds the octet rule?

AlI₃

PO₂⁺

SeCl₂

all of the above

none of the above

Practice: Drawing Best Lewis Structures

Complete the “best” Lewis diagrams for the following species and use them to answer the next three questions



Part 3

Which Lewis diagram has lone pair(s) on the central atom?

AlI_3



PO_2^+



SeCl_2



all of the above



none of the above



[View Solutions on Wizeprep.com](#)

Solutions to these questions, as well as step-by-step breakdowns of the answers at:

3.6 Resonance

3.6.1

Resonance Structures

Resonance structures: are used to describe molecules with **delocalized electrons** which cannot be described by a single Lewis structure.

Resonance structures follow the same rules as Lewis structures and the same features that make a good Lewis structure also make good resonance structures.

WIZE CONCEPT

The atoms will be connected in the exact same way for each resonance structure (no single bonds can change).

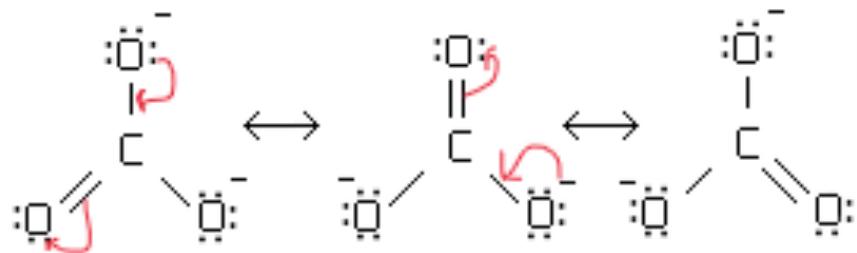
The **only thing that moves are the multiple bonds and lone pairs!**

WIZE TIP

How to Draw Resonance Structures:

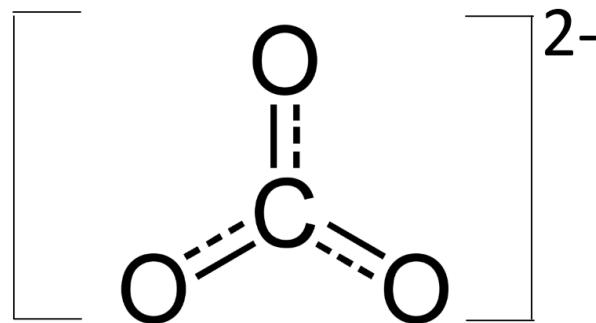
1. If it is not provided, **draw a Lewis structure**.
2. **Move lone pairs and double bond** around to spread out charge.
3. If you move a lone pair into a multiple bond, **make sure your new structure doesn't break the octet rule**.

Example: Resonance structures for CO_3^{2-}



Resonance structures are not discrete molecules which exist individually, but rather the molecule exists as an average of all of its resonance structures.

Resonance hybrid: The "average" of all of a molecule's resonance structures (see below)



Average C-O bond order: $1 \frac{1}{3}$
Average Charge on oxygen: $-2/3$

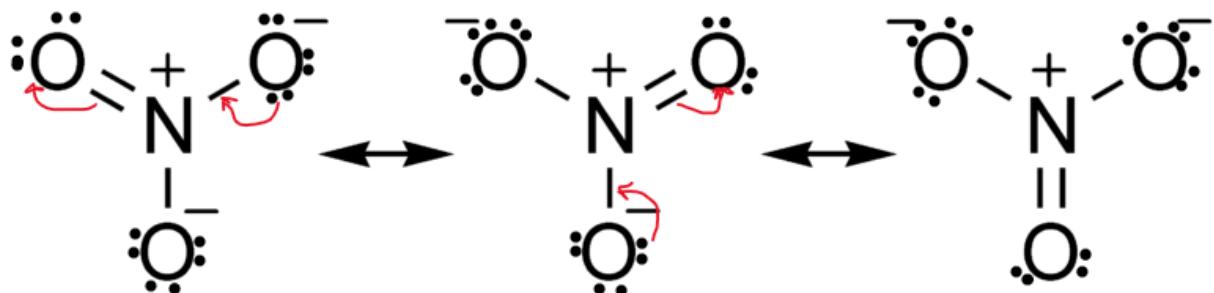
Equivalent Vs Non-Equivalent Resonance Structures

If there is more than one relatively good Lewis structure for a molecule, the molecule exhibits **resonance**.

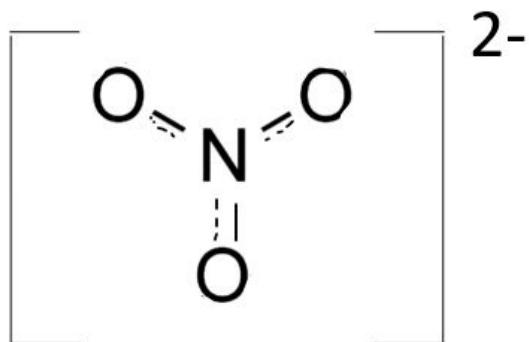
Equivalent Resonance Structures

All resonance structures are **equally stable**.

Example:



In reality, the molecular structure will be the average (hybrid) of each resonance form:



Non-Equivalent Resonance Structures

One resonance structure is better than the other.

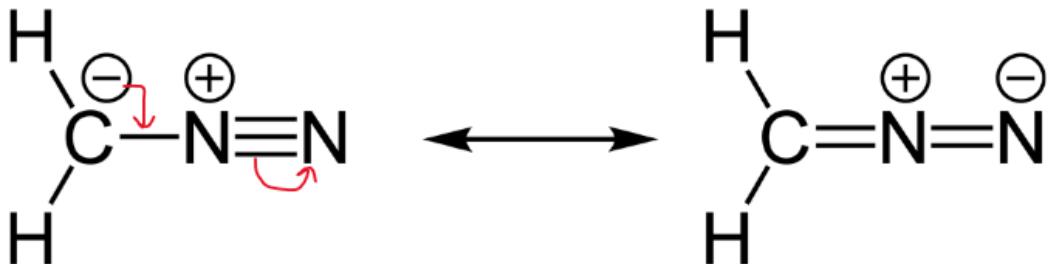
WIZE TIP

The **major contributor** (aka the **more stable resonance structure**) is the one where the **negative charge** resides on the **most electronegative atom**!

The major contributor will also have the **smallest charges** (charges more spread out).

You could be asked to identify the most stable resonance structure on an exam.

Example:



Out of these two non-equivalent resonance structures, which is the more stable one? The one on the left or right? _____

3.6.3

Example: Draw the Resonance Structures

Draw all the reasonable resonance structures for SO_3F^- . Are these equivalent or non-equivalent resonance structures?

Example: Resonance Structures and Stability

Draw all reasonable resonance structures for the following compounds.

Are they equivalent or non-equivalent resonance structures?

If one resonance structure is more stable than the other(s), indicate which one is most stable.

a) N_3^-

b) PO_4^{3-}

3.6.5



MARK YOURSELF QUESTION

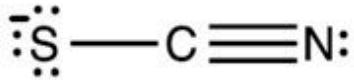
1. Grab a piece of paper and try this problem yourself.
2. When you're done, check the "I have answered this question" box below.
3. View the solution and report whether you got it right or wrong.

Practice: Drawing Arrows for Resonance Structures and Stability

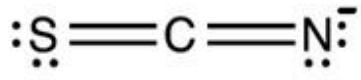
There are three resonance structures one can draw where the octet rule is satisfied for all the atoms of the polyatomic ions SCN^- . These three resonance structures are shown below.

- a) On the Resonance Structures draw curved arrows that show how one could go from Resonance Structure 1 to 2 and Resonance Structure 2 to 3.

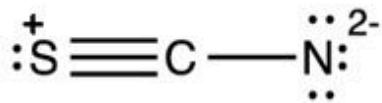
Resonance Structure 1



Resonance Structure 2



Resonance Structure 3



- b) Which of the structure(s) would you identify to be the best Lewis structure(s)? Explain your reasoning.

Practice: Invoking Resonance Structures

For which of the following molecules do we need to invoke resonance to effectively describing the bonding?

