Basic Concepts of Chemical Bonding

Cover 8.1 to 8.7 EXCEPT

1. Omit Energetics of Ionic Bond Formation
   Omit Born-Haber Cycle
2. Omit Dipole Moments

Elements & Compounds

- Why do elements react to form compounds ?
- What are the forces that hold atoms together

  in molecules ?

  and ions in ionic compounds ?

Electron configuration predict reactivity

Element	Electron configurations	
Mg (12e)	$1S^2\ 2S^2\ 2P^6\ 3S^2$	Reactive
Mg^{2+}(10e)	[Ne]	Stable
Cl(17e)	$1S^2\ 2S^2\ 2P^6\ 3S^2\ 3P^5$	Reactive
Cl^- (18e)	[Ar]	Stable

CHEMICAL BONDS

attractive force holding atoms together

Single Bond: involves an electron **pair**

e.g. H_2

Double Bond: involves two electron **pairs**

e.g. O_2

Triple Bond : involves three electron **pairs**

e.g. N_2

Types of Chemical Bonds

Ionic

Polar Covalent

Covalent

Two Extremes

The Two Extremes

IONIC BOND results from the ***transfer*** of electrons from a metal to a nonmetal.

COVALENT BOND results from the **sharing** of electrons between the atoms. Usually found between nonmetals.

The POLAR COVALENT bond
is In-between

- the IONIC BOND [**<u>transfer</u>** of electrons]

and

- the COVALENT BOND [shared electrons]

The pair of electrons in a polar covalent bond are **<u>not shared equally</u>**.

DISCRIPTION OF ELECTRONS

1. How Many Electrons ?
2. Electron Configuration
3. Orbital Diagram
4. *Quantum Numbers*
5. **LEWIS SYMBOLS**

LEWIS SYMBOLS

1. Electrons are represented as **DOTS**
2. Only **VALENCE** electrons are used

Atomic Hydrogen is H •

Atomic Lithium is Li •

Atomic Sodium is Na •

All of Group 1 has only one dot

TABLE 8.1 Lewis Symbols

Element	Electron Configuration	Lewis Symbol	Element	Electron Configuration	Lewis Symbol
Li	[He]$2s^1$	Li·	Na	[Ne]$3s^1$	Na·
Be	[He]$2s^2$	·Be·	Mg	[Ne]$3s^2$	·Mg·
B	[He]$2s^22p^1$	·Ḃ·	Al	[Ne]$3s^23p^1$	·Ȧl·
C	[He]$2s^22p^2$	·Ċ̣·	Si	[Ne]$3s^23p^2$	·Ṡị·
N	[He]$2s^22p^3$	·Ṇ̇:	P	[Ne]$3s^23p^3$	·Ṗ̣:
O	[He]$2s^22p^4$	:Ọ̇:	S	[Ne]$3s^23p^4$	:Ṩ:
F	[He]$2s^22p^5$	·F̤̈:	Cl	[Ne]$3s^23p^5$	·C̤̈l:
Ne	[He]$2s^22p^6$	:N̤̈e:	Ar	[Ne]$3s^23p^6$	:Ä̤r:

The Octet Rule

Atoms gain, lose, or share electrons until they are surrounded by ***8 valence electrons*** ($s^2\ p^6$)

All noble gases [EXCEPT HE] have $s^2\ p^6$ configuration.

Note:

There are exceptions to the octet rule.

I. The Ionic Bond

results from the ***<u>transfer</u>*** of electrons

$$
\begin{array}{ccccc}
 & \cdot\cdot & & & \cdot\cdot \\
\mathrm{Na}\,\bullet \;+ & \bullet\,\mathrm{Cl}\,: & \rightarrow & \mathrm{Na}^{+} & :\mathrm{Cl}:^{-} \\
 & \cdot\cdot & & & \cdot\cdot
\end{array}
$$

Na has lost an electron to become Na^+ and chlorine has gained the electron to become Cl^-

II. Covalent Bonding

results from the **sharing** of electrons between the atoms.

For example

$$H\bullet + \bullet H \rightarrow H\bullet\bullet H \quad \text{or} \quad H:H$$

Each pair of shared electrons constitutes one chemical bond.

Bonding & Non Bonding Electrons

Bonding Electrons: electrons between elements

How many Bonding electrons in

Hydrogen ?   Chlorine ?

NonBonding Electrons: those not used in bonding

How many Non Bonding electrons in

Hydrogen ?   Chlorine ?

Multiple Bonds

$$H-H \qquad \ddot{\underset{\cdot\cdot}{O}}=\ddot{\underset{\cdot\cdot}{O}} \qquad :N\equiv N:$$

One shared pair of electrons single bond

e.g. H_2 H - H

Two shared pairs of electrons double bond

e.g. O_2 ::O = O::

Three shared pairs of electrons triple bond

e.g. N_2 : N ≡ N :

Covalent Bonding

When two atoms of the <u>**same kind**</u> bond, neither of them wants to lose or gain an electron

Therefore, they must share electrons

Each pair of shared electrons constitutes one chemical bond.

Strengths of Covalent Bonds

- We know that multiple bonds are shorter than single bonds.
- We know that multiple bonds are stronger than single bonds.
- As the number of bonds between atoms increases, the atoms are held closer and more tightly together.

III. Polar Covalent Bonds

In a Polar Covalent bond, electrons are shared.

But NOT equal sharing of those electrons.

In Polar Covalent bonds, the electrons are located closer to one atom than the other.

Unequal sharing of electrons results in polar bonds.

Polar Covalent Bonds

$$H^{\delta+} \rightarrow F^{\delta-}$$

There is more electron density on F than on H.

Since there are two different "ends" of the molecule, HF has a **di pole.**

SPECIAL CASE C- H Bond

·

Lewis dot formula · C · **for carbon**

·

H

Lewis dot formula · ·

H · · C · · **H**

· · **for methane**

H

Electronegativity

The ability of one atoms

in a molecule

to attract electrons to itself.

Electronegativity 0.7 (**Cs**) to **4.0** (**F**)

Group 1

H ← EXCEPTION { *HIGH* }

2.1

See Fig 8.6
page 285

Li

1.0

N	**O**	**F**
3.0	3.5	4.0
		Cl
		3.0

Na

0.9

{LOW}

Dipole

The difference in electronegativity leads to a polar covalent bond.

$$H^{\delta+} \rightarrow F^{\delta-}$$

There is more electron density on F than on H.

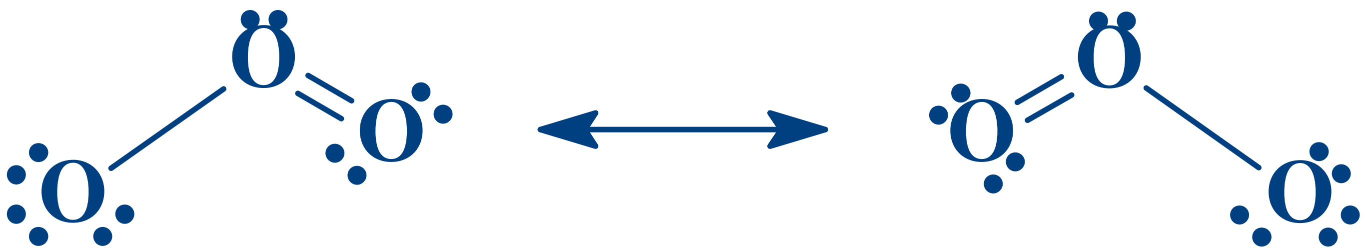
Since there are two different "ends" of the molecule, HF has a **di pole.**

Resonance Structures

Two or more alternative Lewis structures for a molecule.

The inability to described a molecule with a single Lewis stucture.

Resonance in Ozone

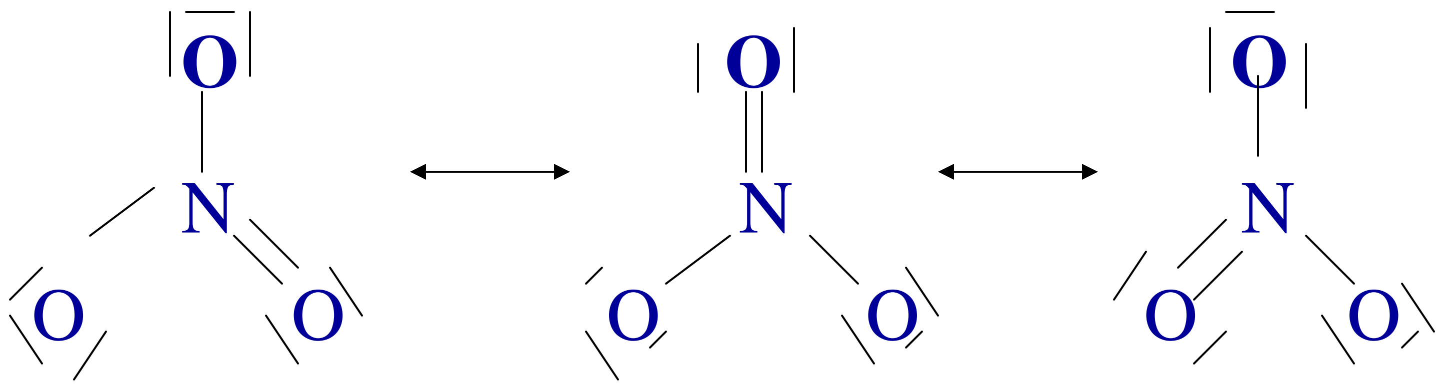


In ozone the extreme possibilities have one double and one single bond.

The resonance structure has two identical bonds of intermediate character

Resonance In Nitrate Ion

In Nitrate Ion **[NO_3^-]** the extreme possibilities have one double and two single bonds



The resonance structure has three identical bonds of intermediate character.

Draw the LEWIS STRUCTURE for

HF

NF_3

H_2O

H_2O_2

$MgCl_2$

CH_4

C_2H_6

Resonance In Nitrite Ion

$(NO_2)^-$

$(O - N - O)^-$

Where does the double bond go ?

$$|\underline{O} = \underline{N} - \overline{\underline{O}}| \longleftrightarrow |\underline{O} - \underline{N} = \overline{\underline{O}}|$$

Formal Charge

The difference between the valence electrons in an isolated atom and the number of electrons assigned to that atom in a Lewis structure.

$$FC = \begin{matrix}\text{number of}\\ \text{valence}\\ \text{electrons}\end{matrix} - \begin{matrix}\text{number of}\\ \text{nonbonding}\\ \text{electrons}\end{matrix} - \frac{1}{2}\begin{matrix}\text{number of}\\ \text{bonding}\\ \text{electrons}\end{matrix}$$

Formal Charge

Example 1

	:: O = C = O::			vs	:::O – C ≡ O		
Valence e^-	6	4	6		6	4	6
- e^- for atom	- 6	- 4	- 6		- 7	- 4	- 5
Formal Charge	0	0	0		-1	0	+1

Correct formula for $(NCO)^{-1}$?

	Structure 1			Structure 2			Structure 3		
	[:::N - C ≡O:]$^{-}$			[::N = C = O::]$^{-}$			[:N ≡ C –O:::]$^{-}$		
V e^-	5	4	6	5	4	6	5	4	6
- e^-	-7	-4	-5	-6	-4	-6	-5	-4	-7
FC	-2	0	+1	-1	0	0	0	0	-1

Structure 3 is correct since negative charge is on the oxygen atom (most electronegative)

Formal Charge

1. Neutral molecules: Formal charges add to zero
2. Ions: Formal charges add to charge on ion
3. Smallest formal charge is preferable
4. Negative formal charge placed on most electronegative element