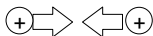


Glencoe Chapter 9

Covalent bonding

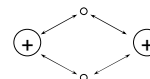
How does H₂ form?

- The nuclei repel



How does H₂ form?

- The nuclei repel
- But they are attracted to electrons
- They share the electrons



Covalent bonds

- Nonmetals hold onto their valence electrons.
- They can't give away electrons to bond.
- Still want noble gas configuration.
- Get it by sharing valence electrons with each other.
- By sharing both atoms get to count the electrons toward noble gas configuration.

Covalent bonding

- Fluorine has seven valence electrons



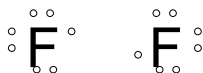
Covalent bonding

- Fluorine has seven valence electrons
- A second atom also has seven



Covalent bonding

- Fluorine has seven valence electrons
- A second atom also has seven
- By sharing electrons



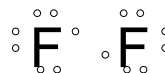
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Covalent bonding

- Fluorine has seven valence electrons
- A second atom also has seven
- By sharing electrons
- Both end with full orbitals

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- A second atom also has seven
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Covalent bonding

- Fluorine has seven valence electrons
- A second atom also has seven
- By sharing electrons
- Both end with full orbitals

Single Covalent Bond

- A sharing of two valence electrons.
- Only nonmetals and Hydrogen.
- Different from an ionic bond because they actually form molecules.
- Two specific atoms are joined.
- In an ionic solid you can't tell which atom the electrons moved from or to.

How to show how they formed

- It's like a jigsaw puzzle.
- I have to tell you what the final formula is.
- You put the pieces together to end up with the right formula.
- For example- show how water is formed with covalent bonds.

Water

Each hydrogen has 1 valence electron
Each hydrogen wants 1 more

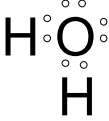
The oxygen has 6 valence electrons
The oxygen wants 2 more
They share to make each other happy

Water

- Put the pieces together
- The first hydrogen is happy
- The oxygen still wants one more

Water

- The second hydrogen attaches
- Every atom has full energy levels



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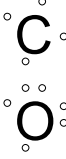
Multiple Bonds

- Sometimes atoms share more than one pair of valence electrons.
- A double bond is when atoms share two pair (4) of electrons.
- A triple bond is when atoms share three pair (6) of electrons.

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Carbon dioxide


- CO₂ - Carbon is central atom (I have to tell you)
- Carbon has 4 valence electrons
- Wants 4 more
- Oxygen has 6 valence electrons
- Wants 2 more



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Carbon dioxide


- Attaching 1 oxygen leaves the oxygen 1 short and the carbon 3 short



22

Carbon dioxide

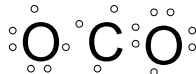
- Attaching the second oxygen leaves both oxygen 1 short and the carbon 2 short



23

Carbon dioxide

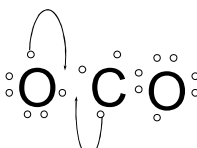
- The only solution is to share more



24

Carbon dioxide

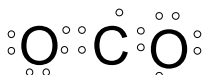
- The only solution is to share more



25

Carbon dioxide

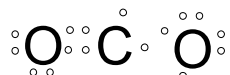
- The only solution is to share more



26

Carbon dioxide

- The only solution is to share more



27

Carbon dioxide

- The only solution is to share more

26

Carbon dioxide

- The only solution is to share more

26

Carbon dioxide

- The only solution is to share more
- Requires two double bonds
- Each atom gets to count all the atoms in the bond

30

Carbon dioxide

- The only solution is to share more
- Requires two double bonds
- Each atom gets to count all the atoms in the bond

31

Carbon dioxide

- The only solution is to share more
- Requires two double bonds
- Each atom gets to count all the atoms in the bond

32

Carbon dioxide

- The only solution is to share more
- Requires two double bonds
- Each atom gets to count all the atoms in the bond

33

How to draw Lewis Structures

- Determine the number of valence electrons for *each* atom.
- Add up the total number of electrons.
- Divide total by 2. This tells you how many electron pairs you have to work with.
- Distribute the electron pairs so that each atom "sees" an octet (full outer shell).

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Examples

- NH_3
- N - has 5 valence electrons
- H - has 1 valence electron each
- NH_3 has $5+3(1) = 8$
- $8/2 = 4$ electron pairs

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Examples

- Draw the electron pairs
- All 8 electrons are accounted for
- Everything is full – each H atom "sees" 2 and the N "sees" 8.

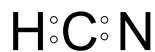
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Examples

- HCN (C is central atom)
- N - has 5 valence electrons
- C - has 4 valence electrons
- H - has 1 valence electron
- HCN has $5+4+1 = 10/2 = 5$ electron pairs

HCN

- Put in single bonds
- Need 2 more bonds
- Must go between C and N



HCN

- Put in single bonds
- Need 2 more bonds
- Must go between C and N
- Uses 8 electrons - 2 more to add



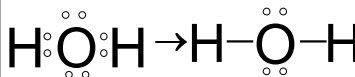
HCN

- Put in single bonds
- Need 2 more bonds
- Must go between C and N
- Uses 8 electrons - 2 more to add
- Must go on N to fill octet



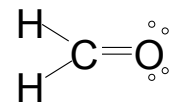
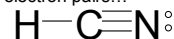
Another way of indicating bonds – Structural Formulas

- Often use a straight line to indicate a pair of shared electrons
- Drawing is called a structural formula



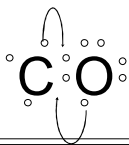
Structural Examples

Be sure to still include non-bonding electron pairs...



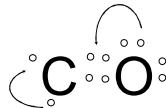
Coordinate Covalent Bond

- When one atom donates both electrons in a covalent bond.
- Carbon monoxide
- CO



Coordinate Covalent Bond

- When one atom donates both electrons in a covalent bond.
- Carbon monoxide
- CO



Coordinate Covalent Bond

- When one atom donates both electrons in a covalent bond.
- Carbon monoxide
- CO



How do we know if coordinate covalent bonding occurs?

- Have to draw the diagram and see what happens.
- Often happens with polyatomic ions and acids.

Resonance

- Resonance occurs when you have the same connections between atoms but the position of a double bond can be on different atoms.
- NO_3^{-1} (common example) – draw the 3 resonant structures here

Polar Bonds

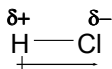
- When the atoms in a bond are the same, the electrons are shared equally.
- This is a nonpolar covalent bond.
- When two different atoms are connected, the atoms may not be shared equally.
- This is a polar covalent bond.
- How do we measure how strong the atoms pull on electrons?

Electronegativity

- A measure of how strongly the atoms attract electrons in a bond.
- The bigger the electronegativity difference the more polar the bond.
- < 0.4 **nonpolar** covalent
- $0.4 \leq 2.0$ **polar** covalent
- > 2.0 ionic – but, just look to see if you have a metal and a non-metal – it's ionic!

How to show a bond is polar

- Isn't a whole charge just a partial charge
- $\delta+$ means a partially positive
- $\delta-$ means a partially negative



- The Cl pulls harder on the electrons
- The electrons spend more time near the Cl

Polar Molecules

"Perfect" **symmetry** – non-polar molecule
vs.

Asymmetry – polar molecule

Polar Molecules

- Molecules with a positive and a negative end
- Requires two things to be true
- * The molecule usually contain polar bonds
This can be determined from differences in electronegativity.
- * Symmetry can not cancel out the effects of the polar bonds.
Must determine geometry first.

Is it polar?

- HF
- H_2O
- NH_3
- CCl_4
- CO_2

Bond Dissociation Energy

- The energy required to break a bond – see chart
- We get the bond dissociation energy back when the atoms are put back together
- If we add up the BDE of the reactants and subtract the BDE of the products we can determine the energy of the reaction (ΔH)
- **Endothermic** – heat needed for reaction to occur (like a "cold" pack)
- **Exothermic** – heat given off as the reaction occurs (like a bomb exploding)

Find the energy change for the reaction

- $\text{CH}_4 + 2\text{O}_2 \longrightarrow \text{CO}_2 + 2\text{H}_2\text{O}$
- For the reactants we need to break 4 C-H bonds at 393 kJ/mol and 2 O=O bonds at 495 kJ/mol = 2562 kJ/mol
- For the products we form 2 C=O at 736 kJ/mol and 4 O-H bonds at 464 kJ/mol
- = 3328 kJ/mol
- reactants - products = 2562-3328 = -766kJ

Intermolecular Forces

What holds molecules to each other

Intermolecular Forces

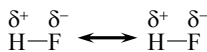
- They are what make solid and liquid molecular compounds possible.
- The weakest are called van der Waal's forces - there are two kinds
- London Dispersion forces - weakest
- Dipole-dipole force of attraction
 - Occurs between molecules that have polar bonds (and are typically polar molecules)

Dipole-dipole forces

- Depend on the number of electrons
- More electrons stronger forces
- Bigger molecules more electrons
 - Fluorine (F_2) is a gas
 - Bromine (Br_2) is a liquid
 - Iodine (I_2) is a solid

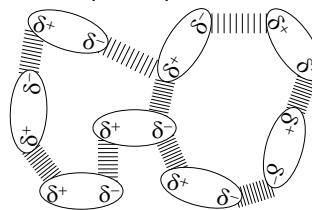
Dipole-dipole forces

- Occur when **polar molecules** are attracted to each other.
- Slightly stronger than dispersion forces.
- Opposites attract but not completely hooked like in ionic solids.



Attraction between partial + end of one molecule to the partial - end of another molecule

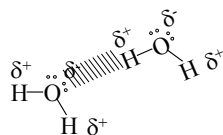
Dipole-dipole forces



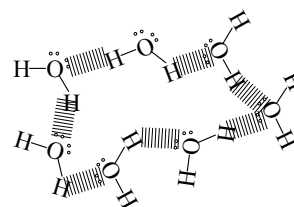
Hydrogen bonding

- The strongest of the intermolecular forces.
- Are the attractive force caused by H bonded to F, O, or N (hence the "hydrogen" bonding)
- F, O, and N are **very electronegative** so it is a **very strong dipole**.
- The hydrogen partially shares with the lone pair e^- in the molecule next to it.

Hydrogen Bonding



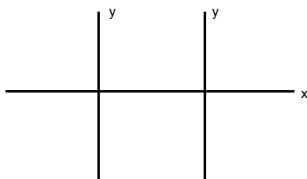
Hydrogen bonding



Sigma and Pi bonds

- Sigma σ - overlapping orbitals "on" an axis
- Pi π - orbitals which overlap but NOT on an axis
- We need to know this information to help us determine molecular shape (which comes up next!)

What do σ bonds and π bonds look like?



Sigma and Pi bonds

- H_2 - two overlapping s orbitals (1 sigma bond)
- HCl - one s orbital and one p orbital overlapping (1 sigma bond)
- O_2 - 1 sigma bond, 1 pi bond
- N_2 - 1 sigma bond, 2 pi bonds

DRAW σ and π bonds on next screen!

