CHEMICAL BONDING

MOLECULAR BONDING

Ball-and-stick model

CHAPTER 8
COVALENT BONDING

SECTION 8.1
MOLECULAR COMPOUNDS

- OBJECTIVES:
  - <u>Distinguish</u> between the melting points and boiling points of *molecular* compounds and *ionic* compounds.
  - <u>Describe</u> the information provided by a *molecular formula*.

BONDS ARE…

✓ Forces that hold groups of atoms together and make them function as a unit. Two types:

1) **Ionic bonds** – *transfer* of electrons (gained or lost; makes **formula unit**)
2) **Covalent bonds** – *sharing* of electrons. The resulting particle is called a "*molecule*"

COVALENT BONDS

- The word ***covalent*** is a combination of the prefix co- (from Latin *com*, meaning "with" or "together"), and the verb *valere*, meaning "to be strong".
- Two electrons shared together have the strength to hold two atoms together in a bond.

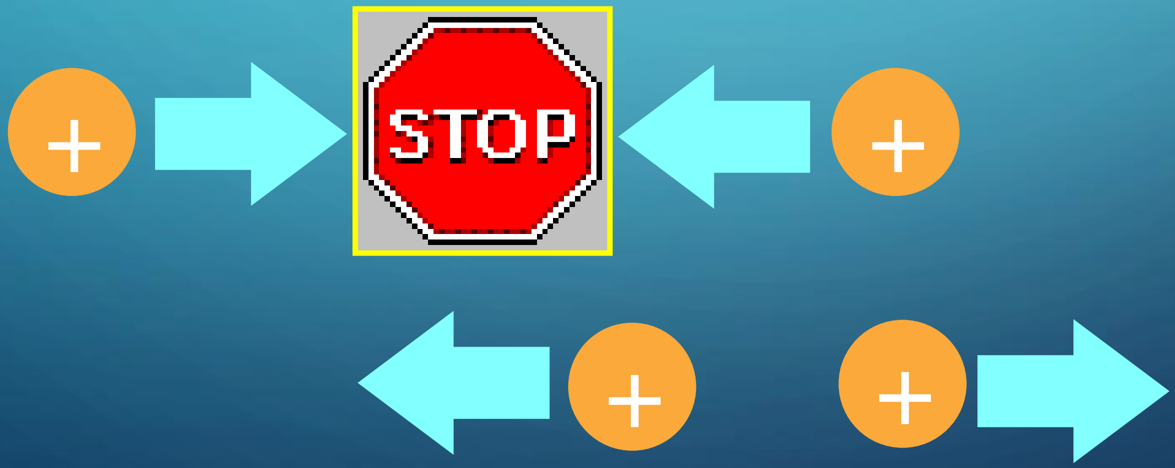
MOLECULES

- Many *elements* found in nature are in the form of *molecules*:
  - a *neutral group of atoms* joined together by covalent bonds.
- For example, air contains oxygen molecules, consisting of two oxygen atoms joined covalently
- Called a "diatomic molecule" (O_2)

HOW DOES H_2 FORM?

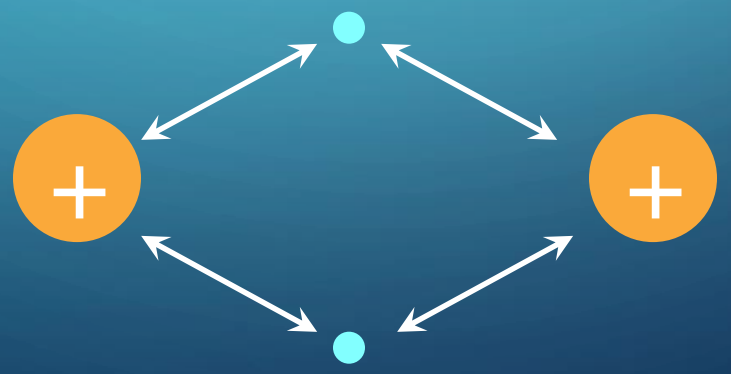
(diatomic hydrogen molecule)

- The nuclei repel each other, since they both have a positive charge (like charges repel).



HOW DOES H_2 FORM?

- But, the nuclei are attracted to the electrons
- They share the electrons, and this is called a "covalent bond", and involves only *NONMETALS*!



CHAPTER 8.1 PRACTICE QUESTIONS

- Page 216 #s 1 & 2

SECTION 8.2
THE NATURE OF COVALENT BONDING

- Describe the different types of covalent bonds
- Explain how co-ordinate covalent bonds are produced
- Explain the role of bond dissociation energy in the properties of a molecule.
- Describe and represent resonance structures.

COVALENT BONDS

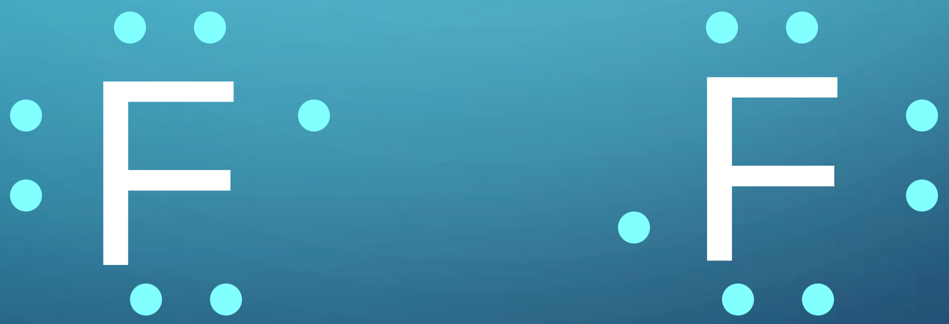
- ***Nonmetals*** hold on to their valence electrons.
- They can't give away electrons to bond.
  - But still want noble gas configuration.
- Get it by sharing valence electrons with each other = *covalent bonding*
- By sharing, **both atoms** get to count the electrons toward a noble gas configuration.

COVALENT BONDING

- Fluorine has seven valence electrons (but would like to have 8)

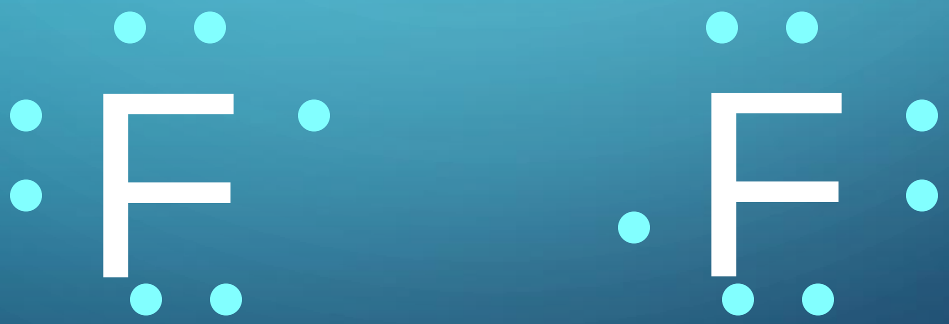
COVALENT BONDING

- 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…



Covalent bonding

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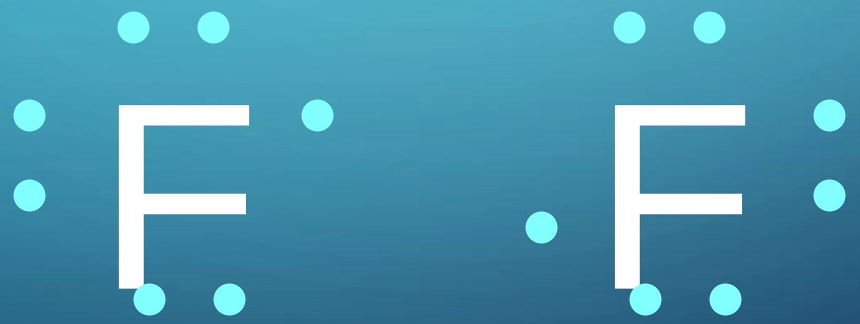

Covalent bonding

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Covalent bonding

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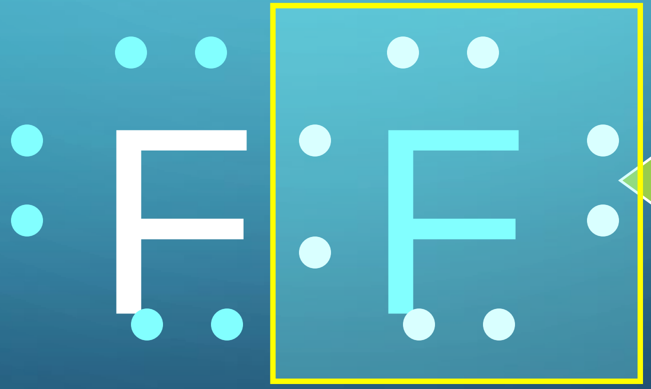
Covalent bonding

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

Covalent bonding

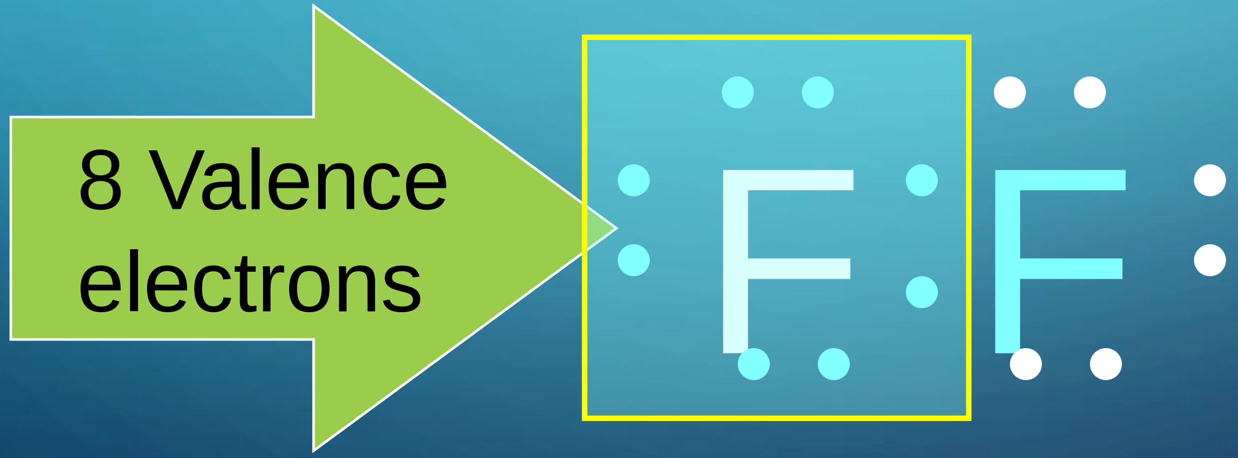
- Fluorine has seven valence electrons
- A second atom also has seven
- By sharing electrons…
- …both end with full orbitals



8 Valence electrons

Covalent bonding

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

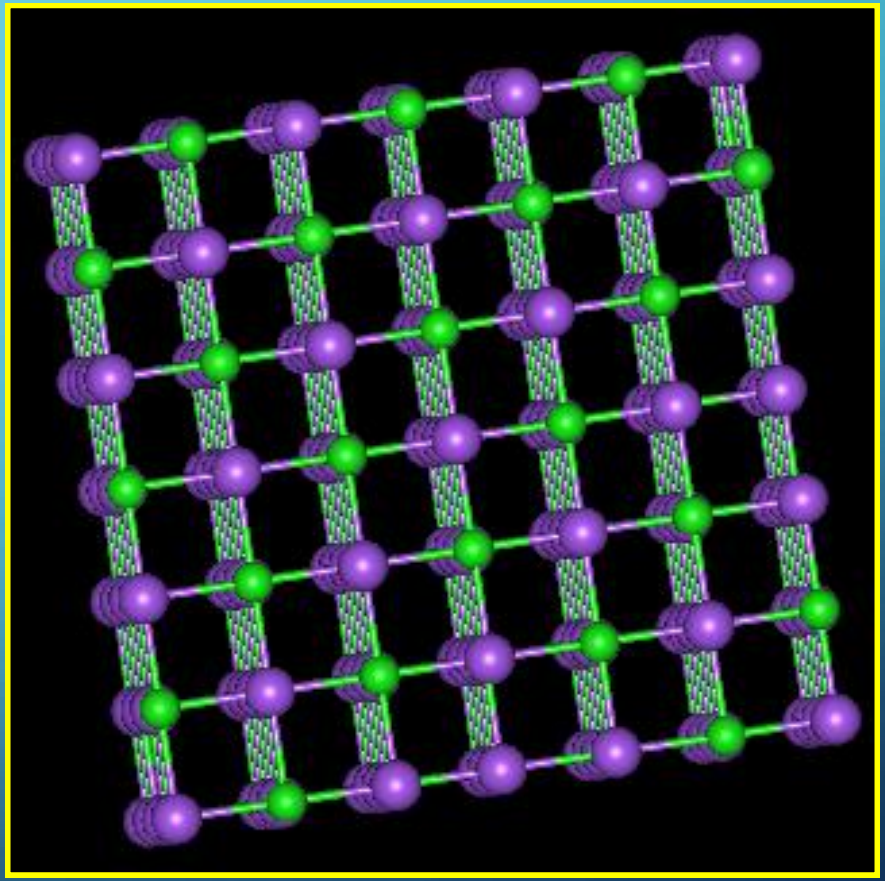


A SINGLE COVALENT BOND IS…

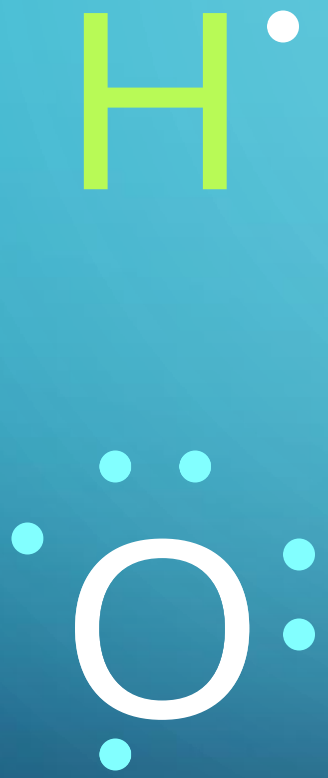
- A sharing of two valence electrons.
- Only nonmetals and hydrogen.
- Different from an ionic bond because they actually form molecules.
- *Two specific atoms* are joined.
- In an ionic solid, you can't tell which atom the electrons moved from or to

SODIUM CHLORIDE CRYSTAL LATTICE

- **Ionic compounds organize in a characteristic *crystal lattice* of alternating positive and negative ions, repeated over and over.**



WATER



✓Each hydrogen has 1 valence electron

- Each hydrogen wants 1 more

✓The oxygen has 6 valence electrons

- The oxygen wants 2 more

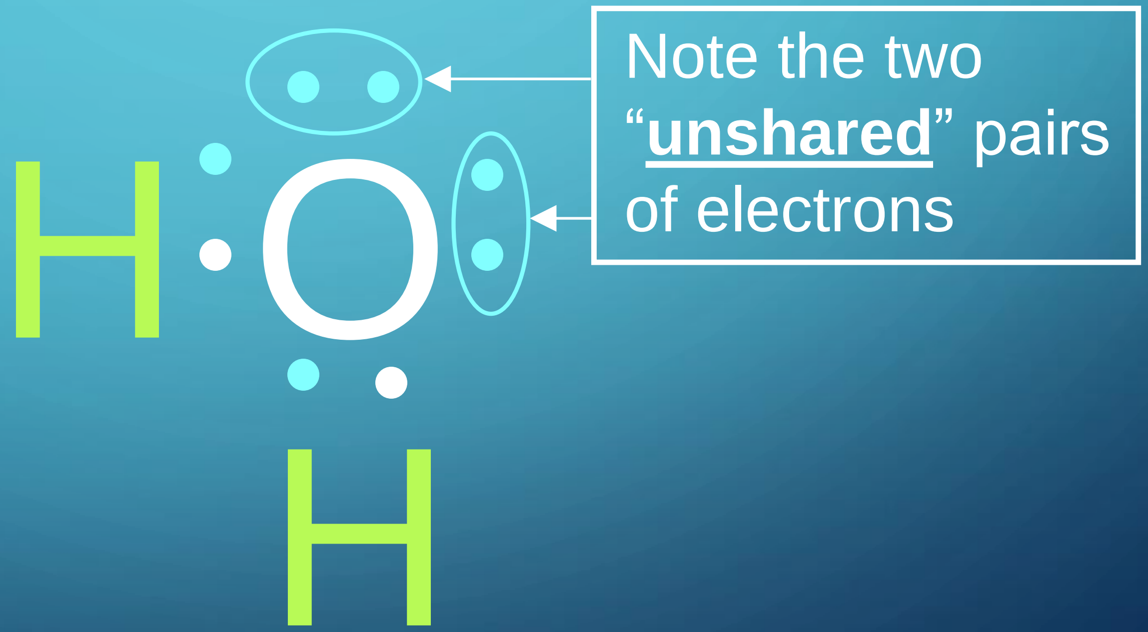
✓They <u>share</u> to make each other complete

WATER

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

WATER

- So, a second hydrogen attaches
- Every atom has full energy levels



Examples:

1. Conceptual Problem 8.1 on page 220
2. Do PCl_3

MULTIPLE BONDS

- Sometimes atoms share *more than one pair* of valence electrons.
- A double bond is when atoms share *two pairs* of electrons (4 total)
- A triple bond is when atoms share *three pairs* of electrons (6 total)
- Table 8.1, p.222 - *Know* these 7 elements as diatomic:

  Br_2 I_2 N_2 Cl_2 H_2 O_2 F_2

What's the deal with the oxygen dot diagram?

DOT DIAGRAM FOR CARBON DIOXIDE

- CO_2 - Carbon is central atom (more metallic)
- Carbon has 4 valence electrons
- Wants 4 more
- Oxygen has 6 valence electrons
- Wants 2 more

CARBON DIOXIDE

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

Carbon dioxide

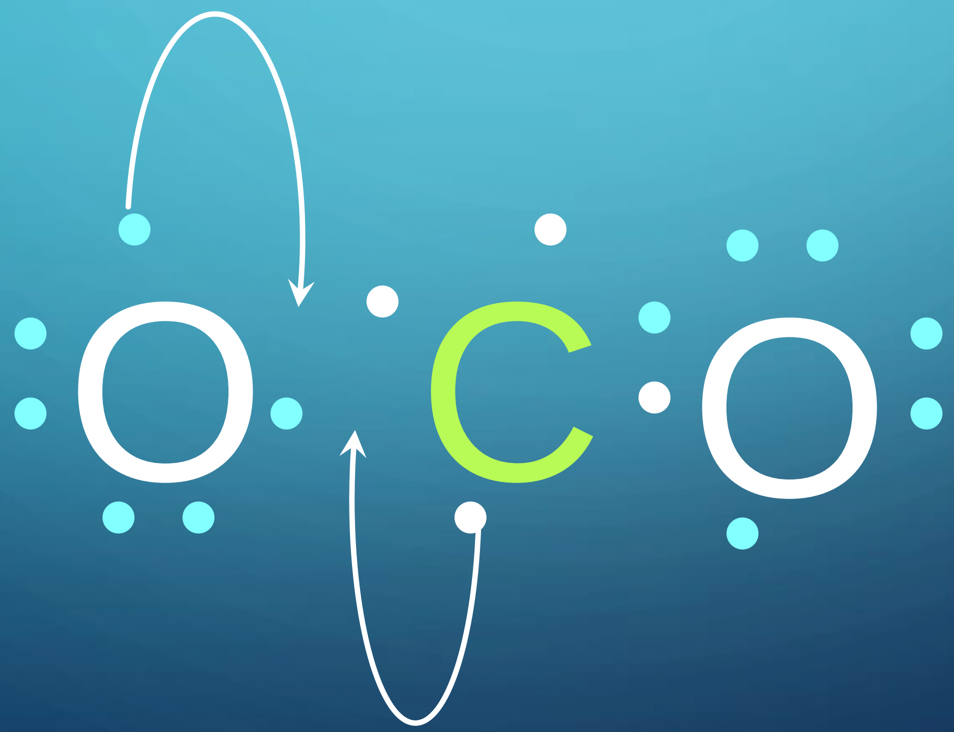
- Attaching the second oxygen leaves both of the oxygen 1 short, and the carbon 2 short

Carbon dioxide

- The only solution is to share more

Carbon dioxide

- The only solution is to share more



Carbon dioxide

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Carbon dioxide

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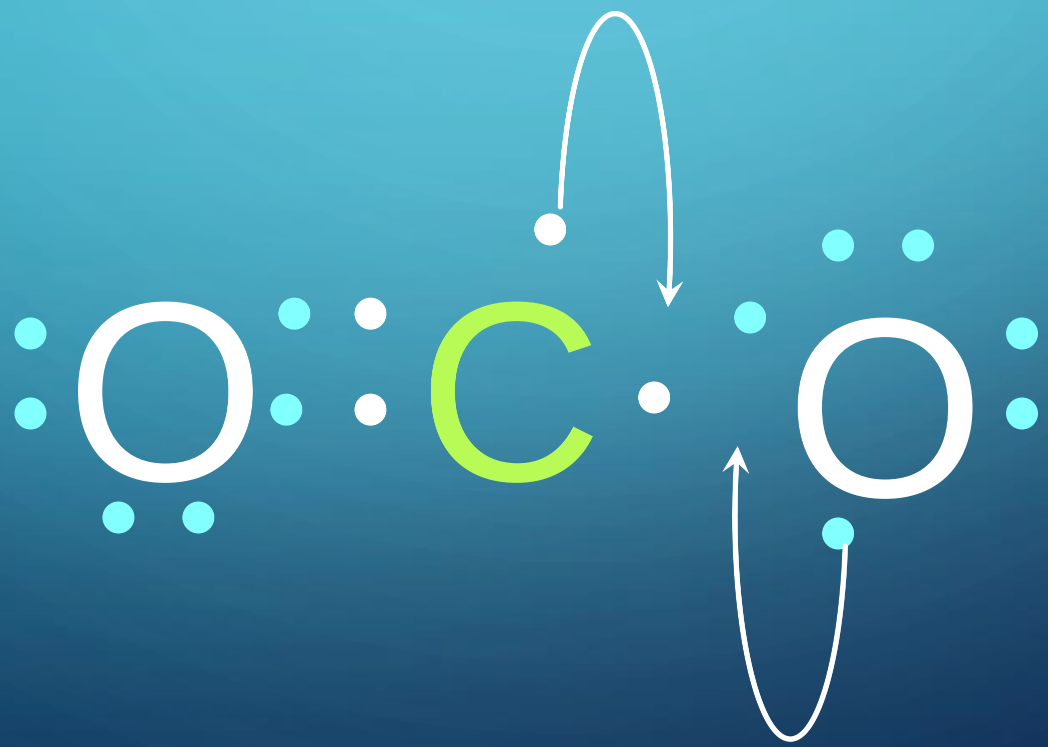
Carbon dioxide

- The only solution is to share more



Carbon dioxide

- The only solution is to share more

Carbon dioxide

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

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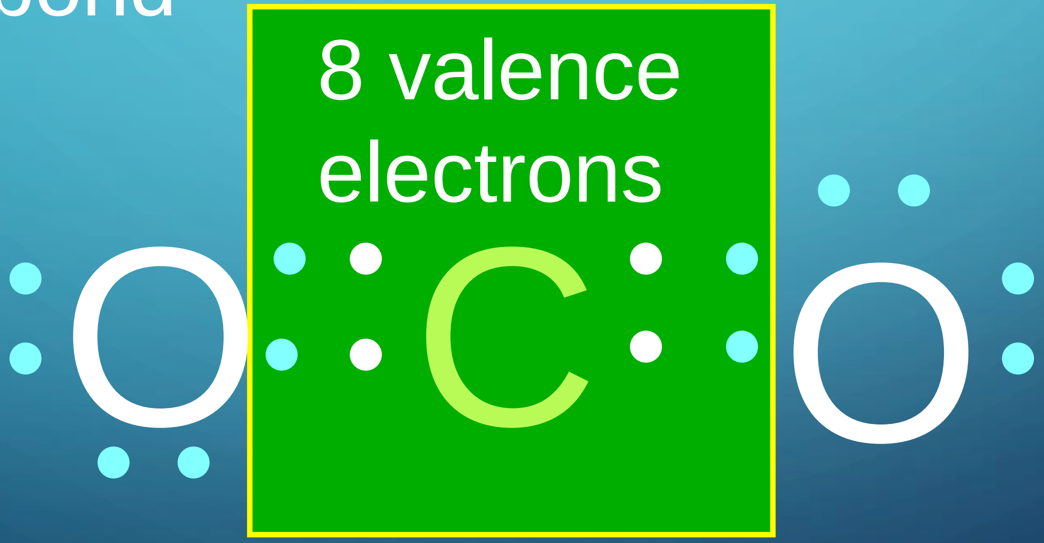
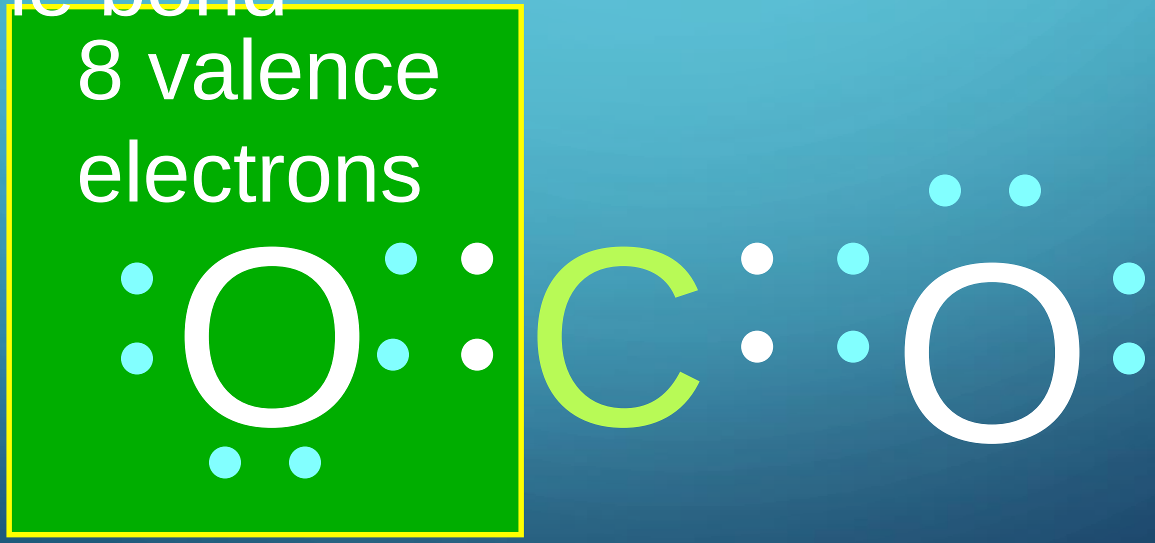


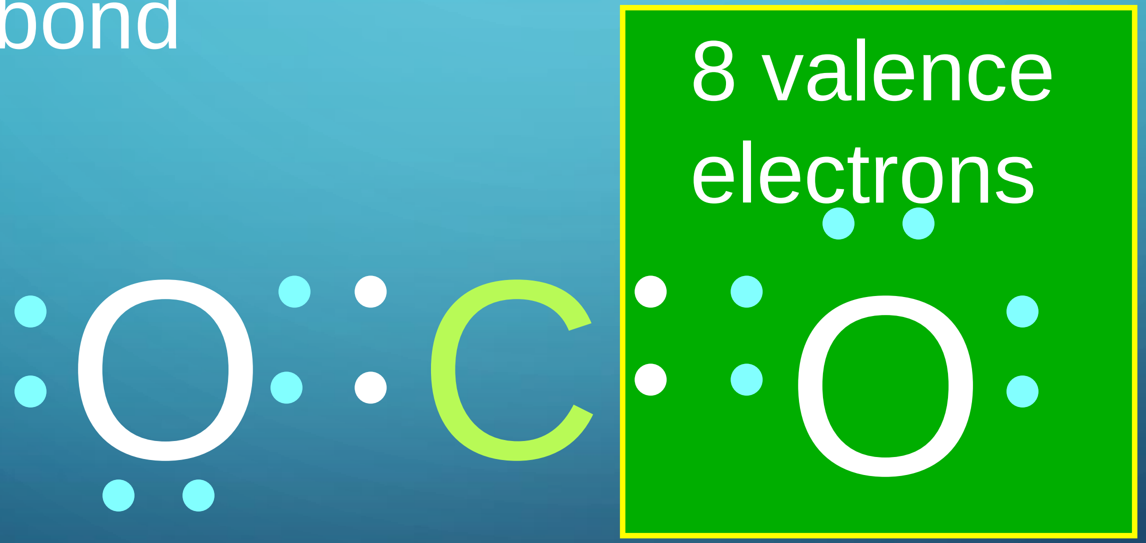
Carbon dioxide

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



Carbon dioxide

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



HOW TO DRAW THEM?

1) Add up all the valence electrons.
2) Count up the total number of electrons to make all atoms happy.
3) Subtract; then Divide by 2
4) Tells you how many bonds to draw
5) Fill in the rest of the valence electrons to fill atoms up.

EXAMPLE

- NH_3, which is ammonia
- N – central atom; has 5 valence electrons, wants 8
- H - has 1 (x3) valence electrons, wants 2 (x3)
- NH_3 <u>has</u> 5+3 = 8
- NH_3 <u>wants</u> 8+6 = 14
- (14-8)/2= 3 bonds
- 4 atoms with 3 bonds

EXAMPLES

- Draw in the bonds; start with singles
- All 8 electrons are accounted for
- Everything is full – done with this one.

EXAMPLE: HCN

- HCN: C is central atom
- N - has 5 valence electrons, wants 8
- C - has 4 valence electrons, wants 8
- H - has 1 valence electron, wants 2
- HCN <u>has</u> 5+4+1 = 10
- HCN <u>wants</u> 8+8+2 = 18
- (18-10)/2= 4 bonds
- 3 atoms with 4 bonds – this will require multiple bonds - not to H however

HCN

- Put single bond between each atom
- Need to add 2 more bonds
- Must go between C and N (Hydrogen is full)

H:C:N

HCN

- Put in single bonds
- Needs 2 more bonds
- Must go between C and N, not the H
- Uses 8 electrons – need 2 more to equal the 10 it has

H:C:::N

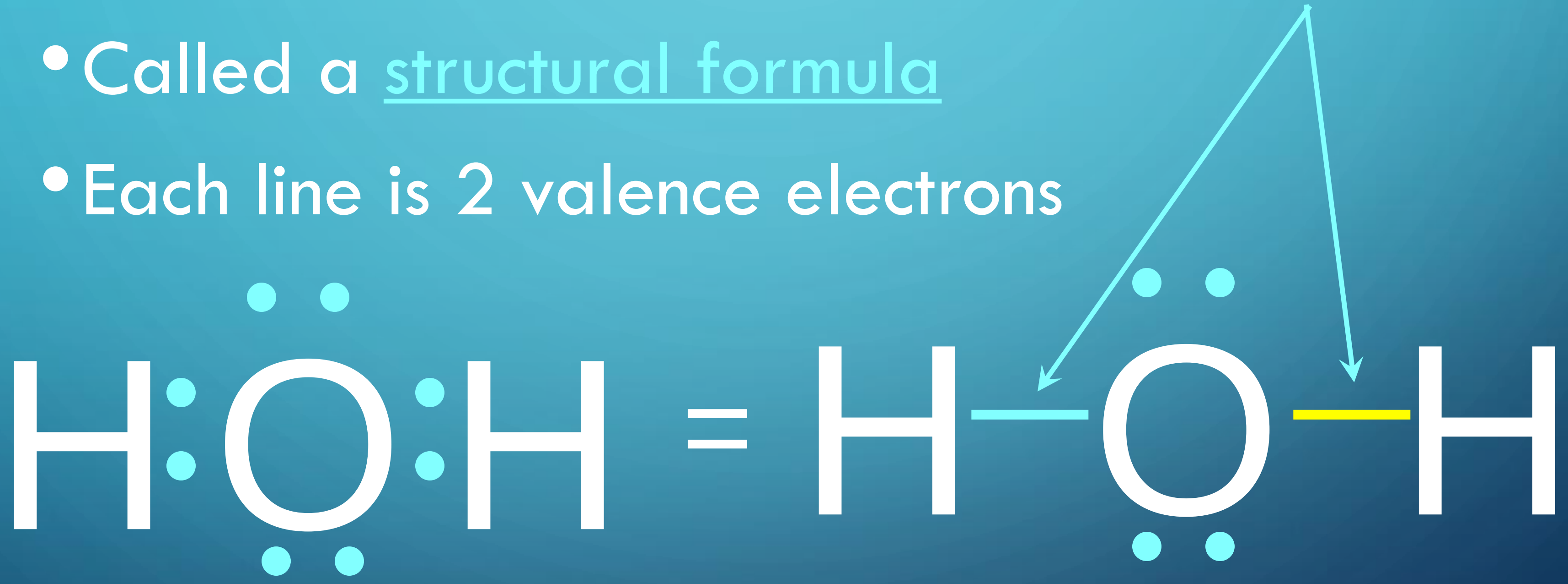
HCN

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

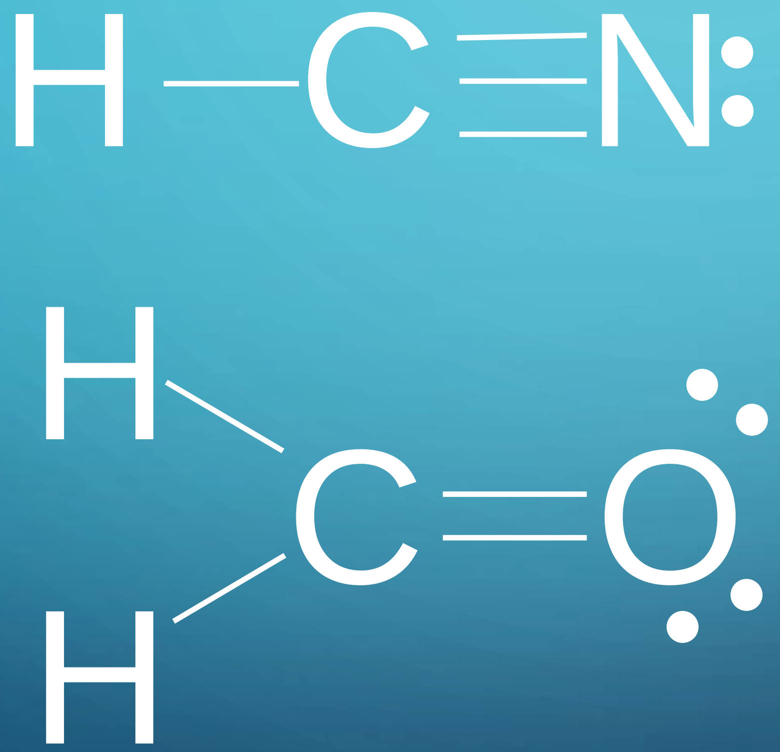
H:C:::N:

ANOTHER WAY OF INDICATING BONDS

- Often use a line to indicate a bond
- Called a structural formula
- Each line is 2 valence electrons



OTHER STRUCTURAL EXAMPLES



A COORDINATE COVALENT BOND…

- When *one atom donates* ***both*** electrons in a covalent bond.
- Carbon monoxide (CO) is a good example:

Both the carbon and oxygen give another single electron to share

Coordinate Covalent Bond

- When one atom donates both electrons in a covalent bond.
- Carbon monoxide (CO) is a good example:

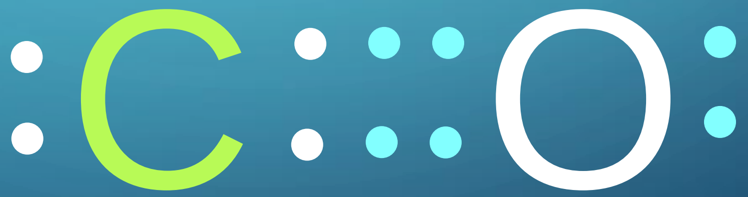
This carbon electron moves to make a pair with the other single.

Oxygen gives both of these electrons, since it has no more singles to share.

Coordinate Covalent Bond

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

The coordinate covalent bond is shown with an arrow as:



COORDINATE COVALENT BOND

- Most *polyatomic cations* and *anions* contain covalent and coordinate covalent bonds
- Table 8.2, p.224
- Conceptual Problem 8.2, p.225
- The ammonium ion (NH_4^{1+}) can be shown as another example

CHAPTER 8.2 PRACTICE QUESTIONS

- Page 229 #s 13 – 16, 20, 21.

SECTION 8.3
BONDING THEORIES

- OBJECTIVES:
  - <u>Describe</u> the relationship between atomic and molecular orbitals.
  - <u>Describe</u> how VSEPR theory helps predict the shapes of molecules.

MOLECULAR ORBITALS ARE…

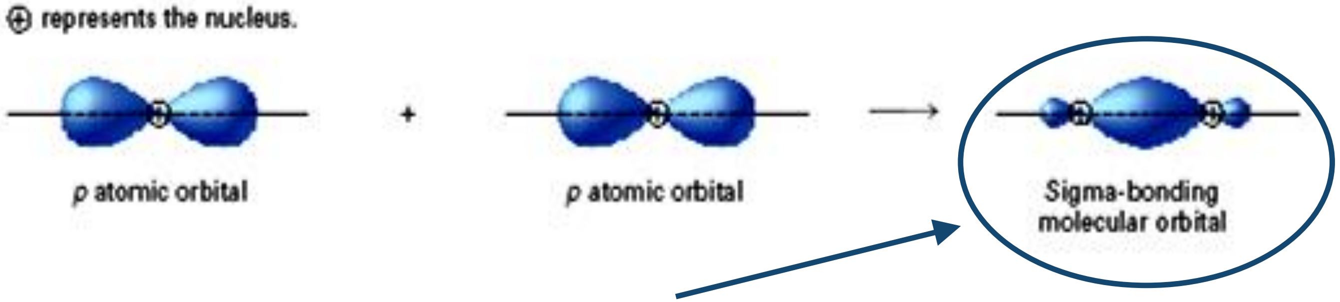
- The model for covalent bonding assumes the orbitals are those of the individual atoms = atomic orbital
- Orbitals that apply to the *overall molecule*, due to atomic orbital ***overlap*** are the molecular orbitals
  - A bonding orbital is a molecular orbital that can be occupied by two electrons of a covalent bond

MOLECULAR ORBITALS - DEFINITIONS

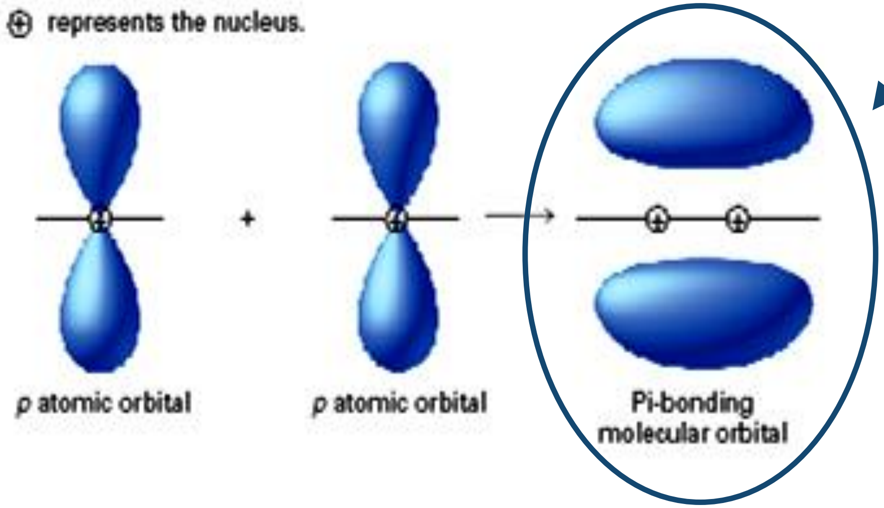
- Sigma bond- when two atomic orbitals combine to form the molecular orbital that is *symmetrical along the axis* connecting the nuclei
- Pi bond- the bonding electrons are likely to be found *above and below* the bond axis (weaker than sigma)
- Note pictures on the next slide

Figures 8.14 and 8.15 Sigma and pi Orbitals

- Pages 230 and 231



Sigma bond is symmetrical along the axis between the two nuclei.



Pi bond is above and below the bond axis, and is weaker than sigma

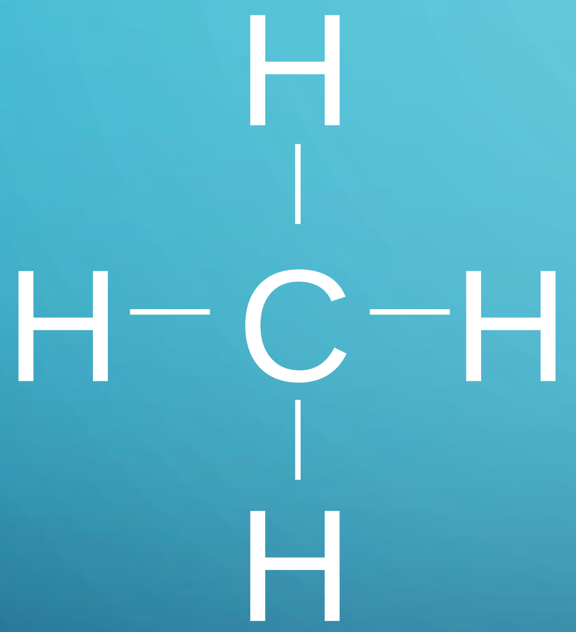
VSEPR: STANDS FOR…

- **V**alence **S**hell **E**lectron **P**air **R**epulsion
- Predicts the three dimensional shape of molecules.
- The name tells you the theory:
  - Valence shell = outside electrons.
  - Electron Pair repulsion = electron pairs try to get as far away as possible from each other.
- Can determine the *angles of bonds*.

VSEPR

- Based on the number of pairs of valence electrons, both *bonded* and *unbonded*.
- Unbonded pair also called lone pair.
- CH_4 - draw the structural formula
- Has $4 + 4(1) = 8$
- wants $8 + 4(2) = 16$
- $(16-8)/2 = 4$ bonds

VSEPR FOR METHANE (A GAS):

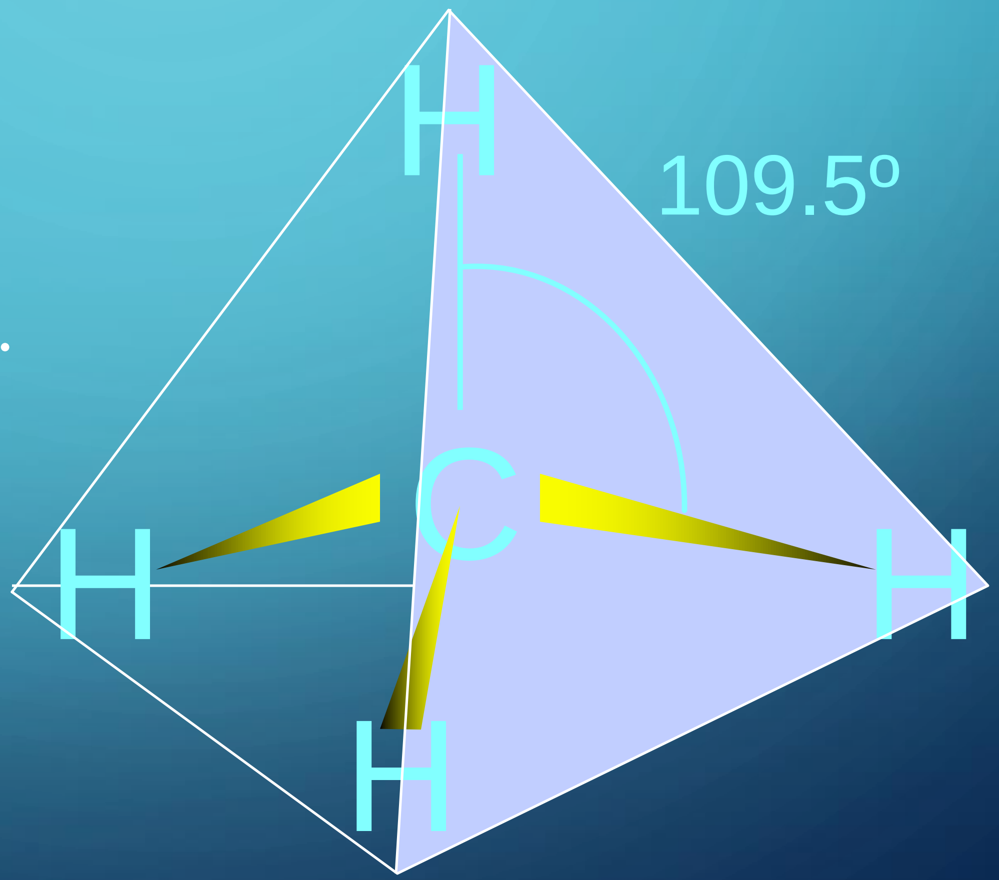


This 2-dimensional drawing does not show a true representation of the chemical arrangement.

- Single bonds fill all atoms.
- There are 4 pairs of electrons pushing away.
- The furthest they can get away is **109.5°**

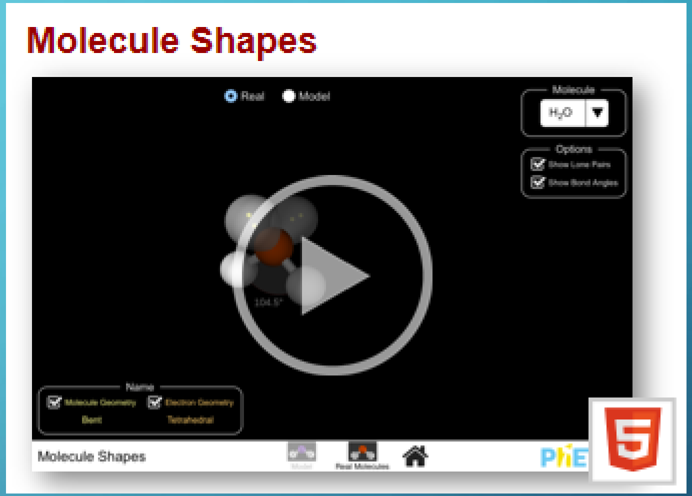
4 ATOMS BONDED

- Basic shape is *tetrahedral*.
- A pyramid with a triangular base.
- Same shape for everything with 4 pairs.



OTHER ANGLES, PAGES 232 - 233

- Ammonia (NH_3) = **107°**
- Water (H_2O) = **105°**
- Carbon dioxide (CO_2) = **180°**

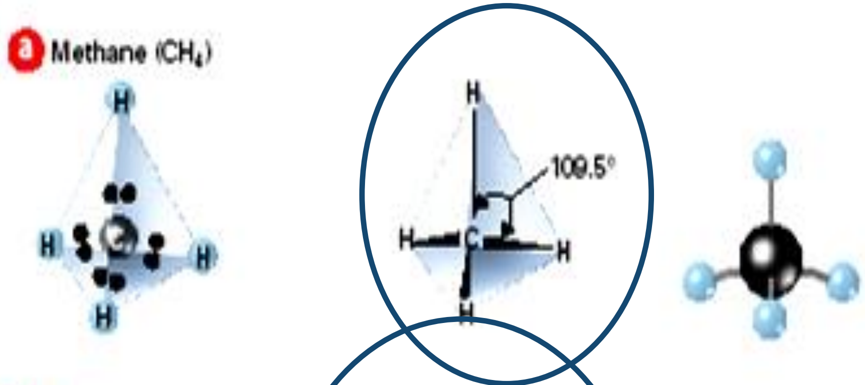


- Note the shapes of these that are pictured on the next slide

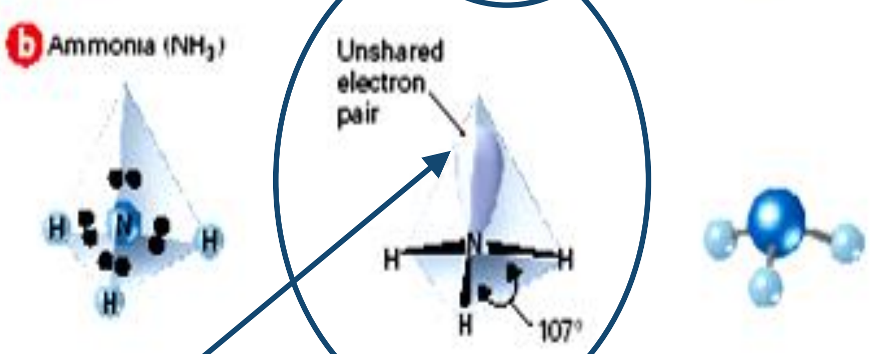
92 Figure 8.16 Tetrahedral and Pyramidal Molecular Structure

- Page 232

Methane has an angle of 109.5º, called **tetrahedral**



Ammonia has an angle of 107º, called **pyramidal**



Note the unshared pair that is repulsion for other electrons.

Chapter 8.3 Practice Questions:

Page 236 #23, 24, 27, 29.

SECTION 8.4
POLAR BONDS AND MOLECULES

- OBJECTIVES:
  - Describe how **electronegativity** values determine the distribution of charge in a polar molecule.
  - Describe what happens to **polar** molecules when they are placed between oppositely charged metal plates.
  - Evaluate the strength of *intermolecular attractions* compared with the strength of ionic and covalent bonds.
  - Identify the reason why *network solids* have high melting points.

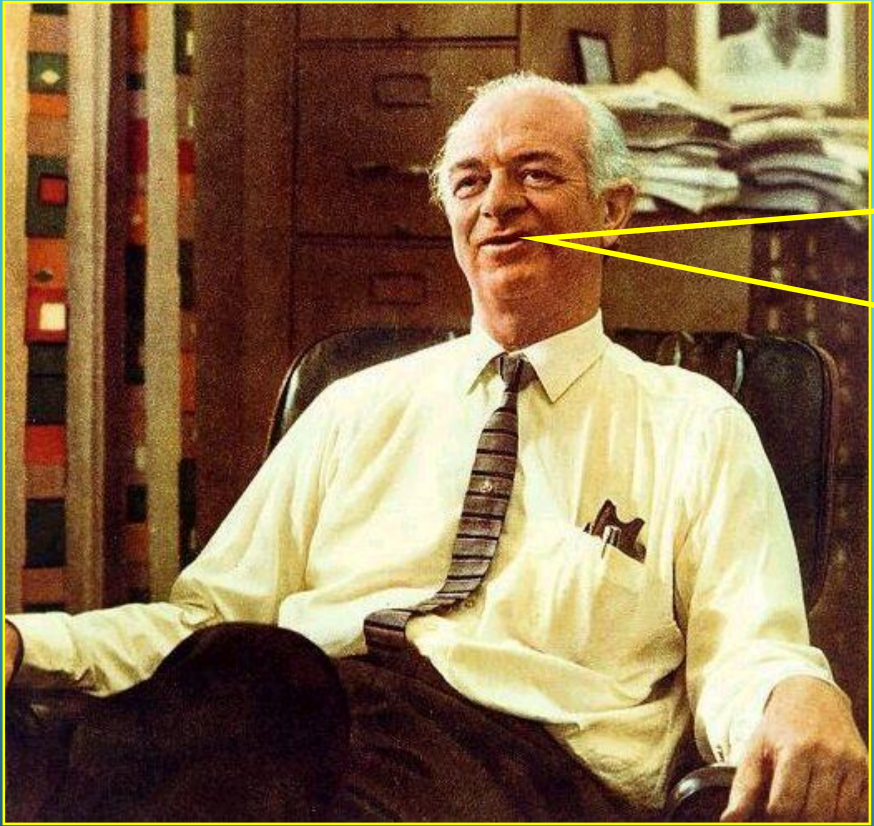
BOND POLARITY

- Covalent bonding means shared electrons
  - but, do they share ***equally***?
- Electrons are pulled, as in a tug-of-war, between the atoms nuclei
  - In **equal** sharing (such as diatomic molecules), the bond that results is called a nonpolar covalent bond

BOND POLARITY

- When two <u>*different* atoms</u> bond covalently, there is an <u>***unequal*** *sharing*</u>
  - the *more electronegative atom* will have a stronger attraction, and will acquire a slightly negative charge
  - called a <u>polar covalent bond</u>, or simply **<u>polar bond</u>**.

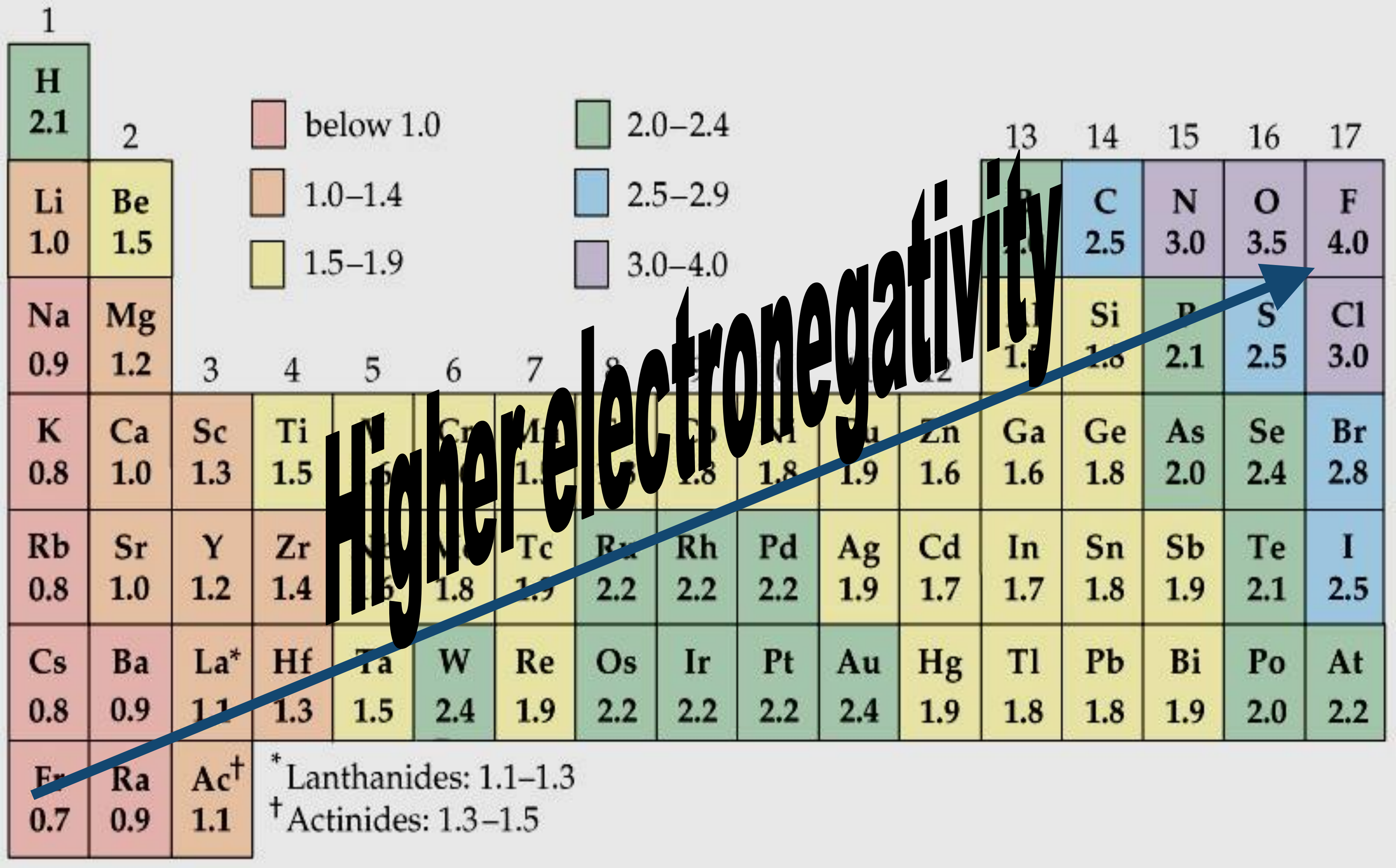
ELECTRONEGATIVITY?



Linus Pauling
1901 - 1994

The ability of an atom in a molecule to attract shared electrons to itself.

TABLE OF ELECTRONEGATIVITIES



1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17
H 2.1																
Li 1.0	Be 1.5											[illegible]	C 2.5	N 3.0	O 3.5	F 4.0
Na 0.9	Mg 1.2											[illegible]	Si 1.8	P 2.1	S 2.5	Cl 3.0
K 0.8	Ca 1.0	Sc 1.3	Ti 1.5	[illegible]	[illegible]	[illegible]	[illegible]	[illegible]	Ni 1.8	Cu 1.9	Zn 1.6	Ga 1.6	Ge 1.8	As 2.0	Se 2.4	Br 2.8
Rb 0.8	Sr 1.0	Y 1.2	Zr 1.4	[illegible]	Mo 1.8	Tc 1.9	Ru 2.2	Rh 2.2	Pd 2.2	Ag 1.9	Cd 1.7	In 1.7	Sn 1.8	Sb 1.9	Te 2.1	I 2.5
Cs 0.8	Ba 0.9	La* 1.1	Hf 1.3	Ta 1.5	W 2.4	Re 1.9	Os 2.2	Ir 2.2	Pt 2.2	Au 2.4	Hg 1.9	Tl 1.8	Pb 1.8	Bi 1.9	Po 2.0	At 2.2
Fr 0.7	Ra 0.9	Ac† 1.1														

*Lanthanides: 1.1–1.3
†Actinides: 1.3–1.5

BOND POLARITY

- Refer to Table 6.2, p. 177.
- Consider HCl

  H = electronegativity of 2.1

  Cl = electronegativity of 3.0

  - the bond is polar
  - the chlorine acquires a slight negative charge, and the hydrogen a slight positive charge

BOND POLARITY

- *Only partial charges*, much less than a true 1+ or 1- as in ionic bond
- Written as:

$$\overset{\delta^{+}}{H}—\overset{\delta^{-}}{Cl}$$

- the positive and minus signs (with the lower case delta: δ^{+} and δ^{-}) denote partial charges.

BOND POLARITY

- Can also be shown:

$$\overset{+\!\!\!\longrightarrow}{\mathrm{H - Cl}}$$

  - the arrow points to the more electronegative atom.

✓ <u>Table 8.3, p.238</u> shows how the electronegativity can also indicate the type of bond that tends to form

Periodic Table of Electronegativities

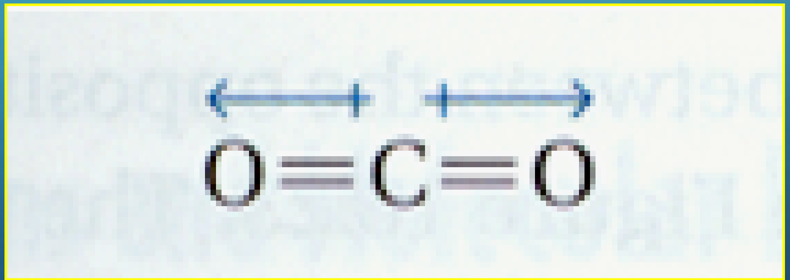
H 2.1																	He
Li 1.0	Be 1.5											B 2.0	C 2.5	N 3.0	O 3.5	F 4.0	Ne
Na 0.9	Mg 1.2											Al 1.5	Si 1.8	P 2.1	S 2.5	Cl 3.0	Ar
K 0.8	Ca 1.0	Sc 1.3	Ti 1.5	V 1.6	Cr 1.6	Mn 1.5	Fe 1.8	Co 1.9	Ni 1.8	Cu 1.9	Zn 1.6	Ga 1.6	Ge 1.8	As 2.0	Se 2.4	Br 2.8	Kr
Rb 0.8	Sr 1.0	Y 1.2	Zr 1.4	Nb 1.6	Mo 1.8	Tc 1.9	Ru 2.2	Rh 2.2	Pd 2.2	Ag 1.9	Cd 1.7	In 1.7	Sn 1.8	Sb 1.9	Te 2.1	I 2.5	Xe
Cs 0.7	Ba 0.9	Lu	Hf 1.3	Ta 1.5	W 1.7	Re 1.9	Os 2.2	Ir 2.2	Pt 2.2	Au 2.4	Hg 1.9	Tl 1.8	Pb 1.9	Bi 1.9	Po 2.0	At 2.2	Rn
Fr 0.7	Ra 0.9	Lr	Rf	Db	Sg	Bh	Hs	Mt	Ds	Uuu	Uub	Uut	Uuq	Uup	Uuh	Uus	Uuo

POLAR MOLECULES

- Conceptual Problem 8.3, p.239
- A polar bond <u>tends</u> to make the **entire molecule** "polar"
  - areas of "difference"
- HCl has polar bonds, thus is a polar molecule.
  - A molecule that has two poles is called <u>dipole</u>, like HCl

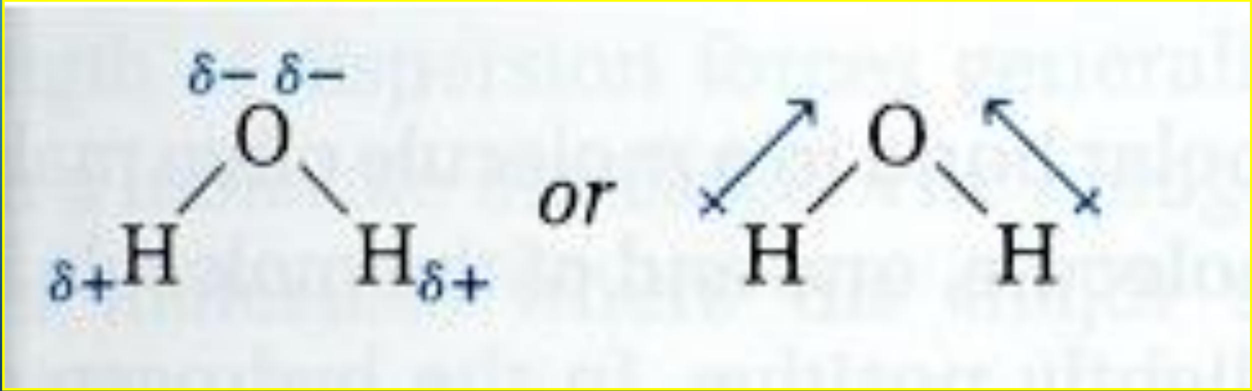
POLAR MOLECULES

- The effect of polar bonds on the polarity of the entire molecule depends on the <u>molecule shape</u>
  - carbon dioxide has two polar bonds, and is linear = nonpolar molecule!



POLAR MOLECULES

- The effect of polar bonds on the polarity of the entire molecule depends on the molecule shape
  - water has two polar bonds and a bent shape; the highly electronegative oxygen pulls the e^- away from H = very polar!



POLAR MOLECULES

- When polar molecules are placed between oppositely charged plates, they tend to become oriented with respect to the positive and negative plates.
- Figure 8.24, page 239

ATTRACTIONS BETWEEN MOLECULES

- They are what make solid and liquid molecular compounds possible.
- The weakest are called van der Waal's forces - there are two kinds:

#1. Dispersion forces

weakest of all, caused by motion of e^-

increases as # e^- increases

halogens start as gases; bromine is liquid; iodine is solid – all in Group 7A

#2. DIPOLE INTERACTIONS

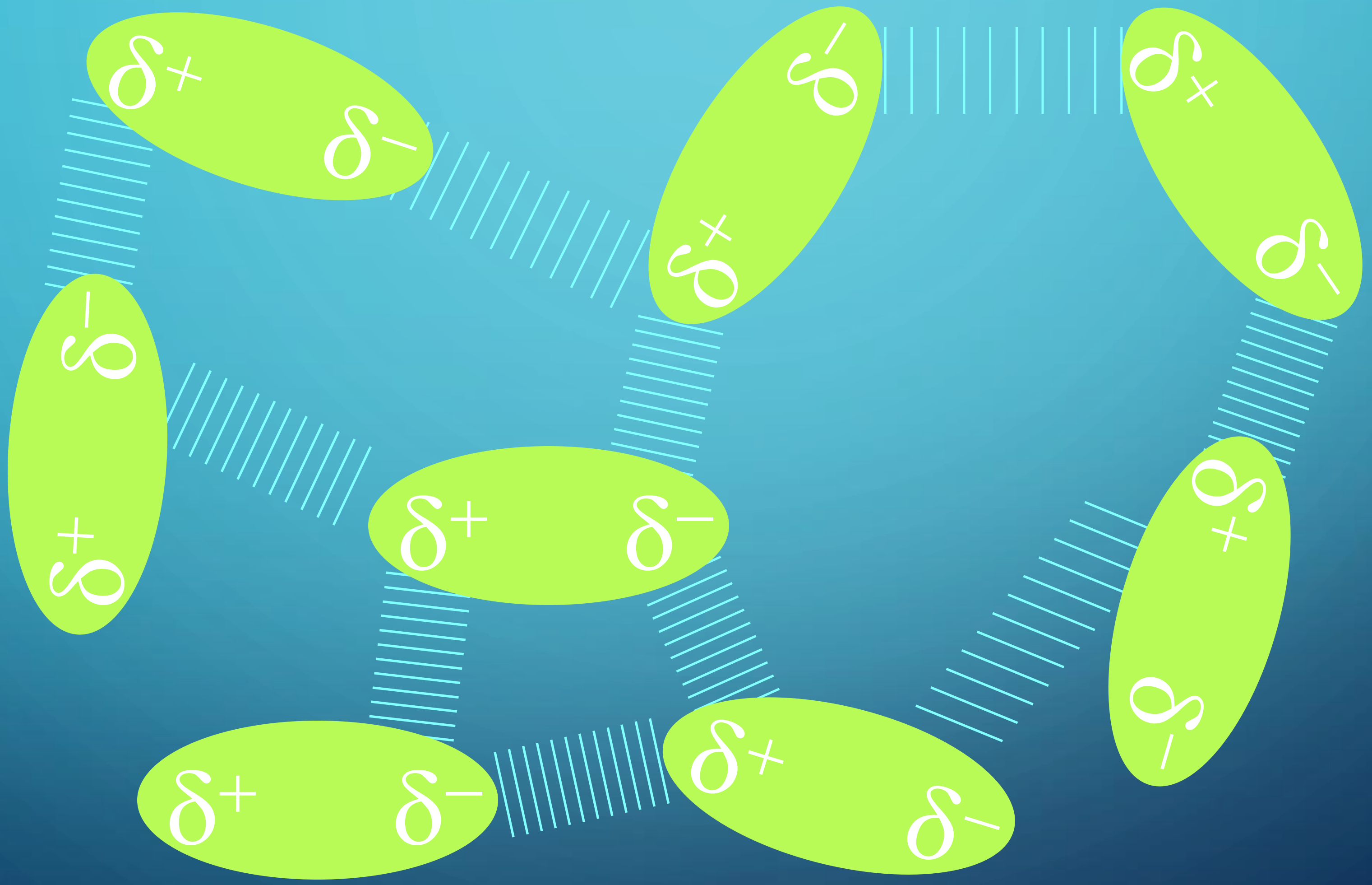
- Occurs when polar molecules are attracted to each other.
- 2. Dipole interaction happens in water
    - Figure 8.25, page 240
    - positive region of one molecule attracts the negative region of another molecule.

#2. DIPOLE INTERACTIONS

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

$$\overset{\delta^+}{\mathrm{H}}\text{—}\overset{\delta^-}{\mathrm{F}} \longleftrightarrow \overset{\delta^+}{\mathrm{H}}\text{—}\overset{\delta^-}{\mathrm{F}}$$

#2. DIPOLE INTERACTIONS



#3. HYDROGEN BONDING

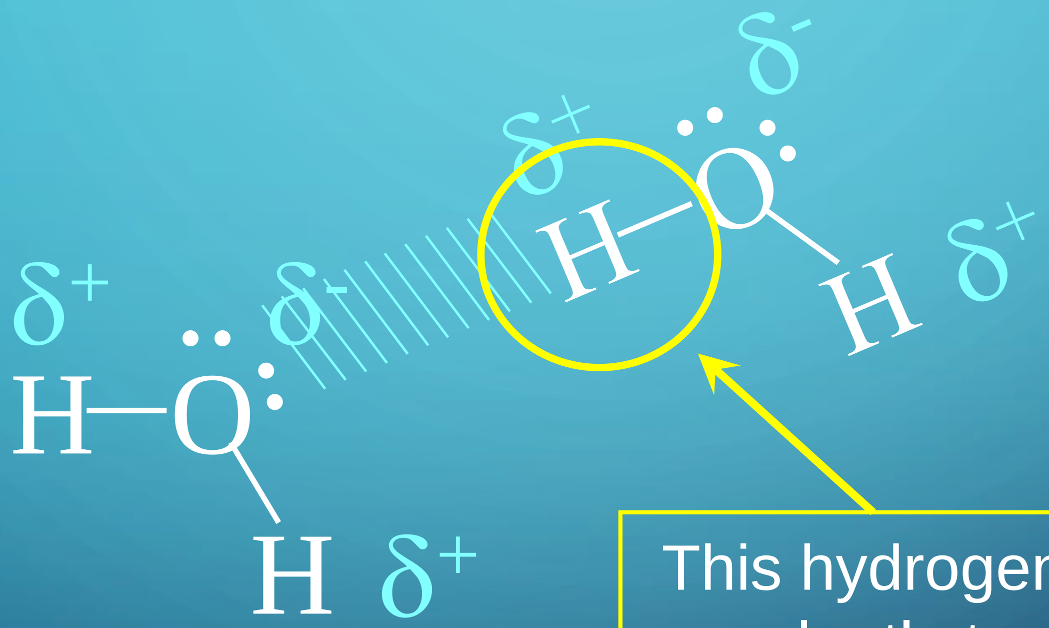
- …is the attractive force caused by hydrogen bonded to N, O, F, or Cl
- N, O, F, and Cl are <u>very</u> electronegative, so this is a very strong dipole.
- And, the hydrogen shares with the lone pair in the molecule <u>*next to it*</u>.
- This is the **strongest** of the intermolecular forces.

#3. HYDROGEN BONDING DEFINED:

- When a hydrogen atom is: a) covalently bonded to a *highly electronegative* atom, AND b) is also weakly bonded to an *unshared electron pair* of a nearby highly electronegative atom.
  - The hydrogen is left very *electron deficient* (*it only had 1 to start with!*) thus it shares with something nearby
  - Hydrogen is also the ONLY element with no shielding for its nucleus when involved in a covalent bond!

HYDROGEN BONDING

(SHOWN IN WATER)



This hydrogen is bonded covalently to: 1) the highly negative oxygen, and 2) a nearby unshared pair.

HYDROGEN BONDING ALLOWS H_2O TO BE A LIQUID AT ROOM CONDITIONS.

ORDER OF INTERMOLECULAR ATTRACTION STRENGTHS

1) Dispersion forces are the weakest
2) A little stronger are the dipole interactions
3) The strongest is the hydrogen bonding
4) ***<u>All of these are weaker than ionic bonds</u>***

ATTRACTIONS AND PROPERTIES

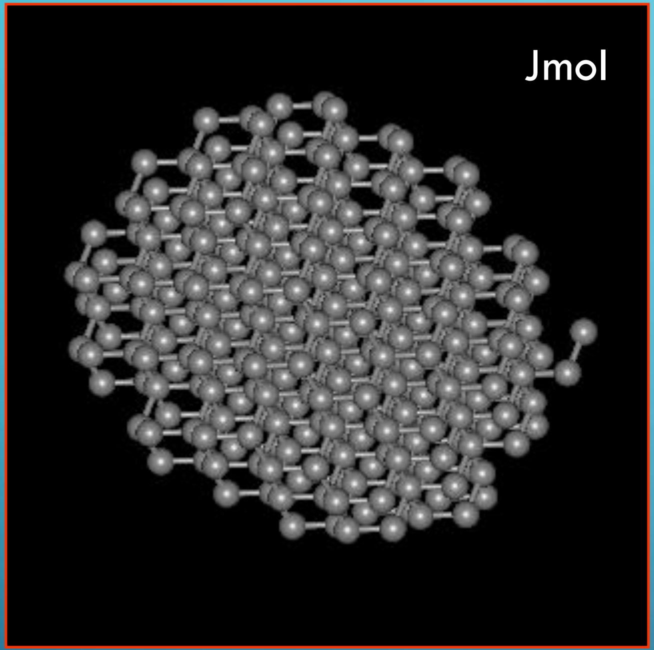
- Why are some chemicals gases, some liquids, some solids?
    - Depends on the type of bonding!
    - Table 8.4, page 244
- Network solids – solids in which ***all the atoms*** *are covalently bonded to* ***each other***

ATTRACTIONS AND PROPERTIES

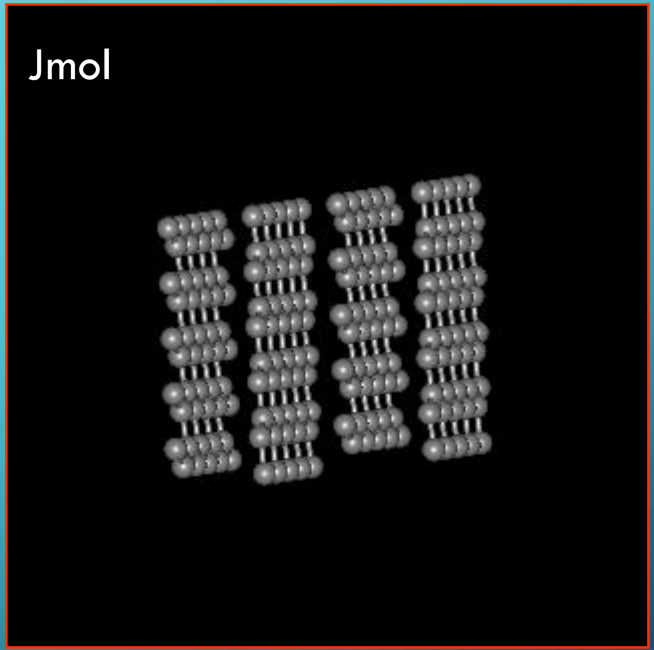
- Figure 8.28, page 243
- Network solids melt at very high temperatures, or not at all (decomposes)
  - Diamond does not really melt, but vaporizes to a gas at 3500 °C and beyond
  - SiC, used in grinding, has a melting point of about 2700 °C

COVALENT NETWORK COMPOUNDS

Some covalently bonded substances DO NOT form discrete molecules.



Diamond, a network of covalently bonded carbon atoms



Graphite, a network of covalently bonded carbon atoms

CHAPTER 8.4 PRACTICE QUESTIONS

- 8.4 Section Assessment #s 32 - 38

QUESTIONS FOR REVIEW

- The following will be organized according to the learning targets as we get closer to the end of the unit:
- Page 247 – 250 #39, 40, 41, 42, 43, 44, 45, 46, 47, 50, 52, 58, 59, 63, 64, 69, 81, 82, 85, 86, 87, 89, 91, 92, 95, 96, 98, 99, 100, 101, 102