

COVALENT BONDING

Unit 1
Chapter 2

The Nature of Covalent Bonding

▣ OBJECTIVES:

- Use electron dot structures to show the formation of single, double, and triple covalent bonds.

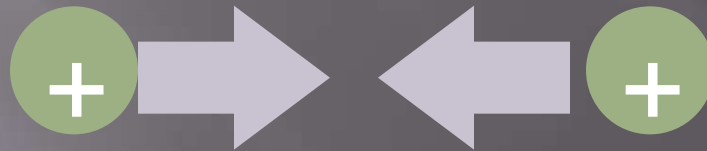
The Nature of Covalent Bonding

▣ OBJECTIVES:

- ▣ Describe and give examples of coordinate covalent bonding, resonance structures, and exceptions to the octet rule.

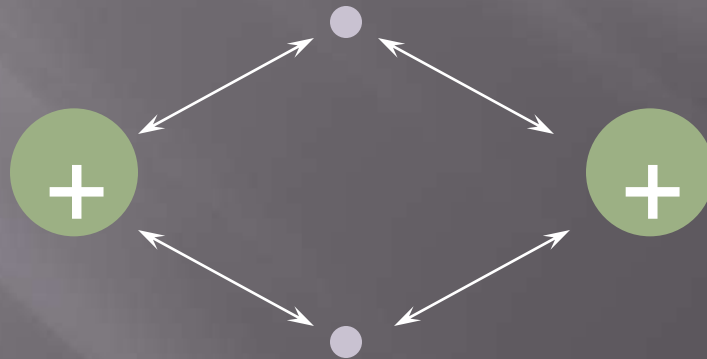
How does H₂ form?

- ▣ The nuclei repel



How does H₂ form?

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



Covalent bonds

- ▣ Nonmetals hold on to 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 a 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

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- A second atom also has seven
- By sharing electrons...



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Covalent bonding

- Fluorine has seven valence electrons
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- By sharing electrons...
- ...both end with full orbitals



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8 Valence
electrons

Covalent bonding

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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.

How to show how they formed

- ▣ It's like a jigsaw puzzle.
- ▣ You put the pieces together to end up with the right formula.
- ▣ Carbon is a special example - can it really share 4 electrons?
 - Electron promotion!
- ▣ Another 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

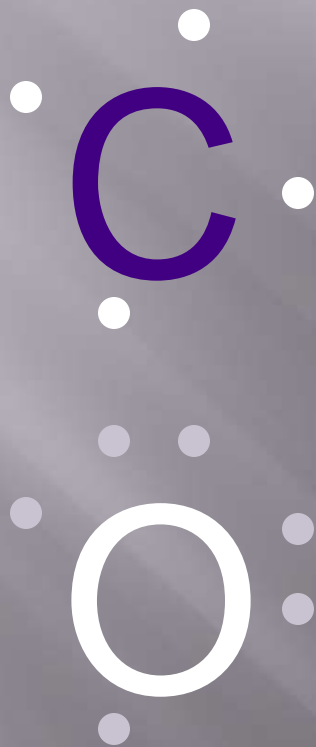
- ▣ The second hydrogen attaches
- ▣ Every atom has full energy levels



Multiple Bonds

- ▣ Sometimes atoms share more than one pair of valence electrons.
- ▣ A double bond is when atoms share two pairs (4 total) of electrons
- ▣ A triple bond is when atoms share three pairs (6 total) of electrons
- ▣ Table 16.1, p.443 - Know which elements are diatomic (Oxygen?)

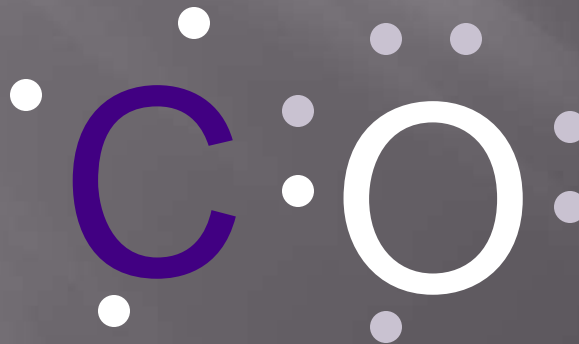
Carbon dioxide



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

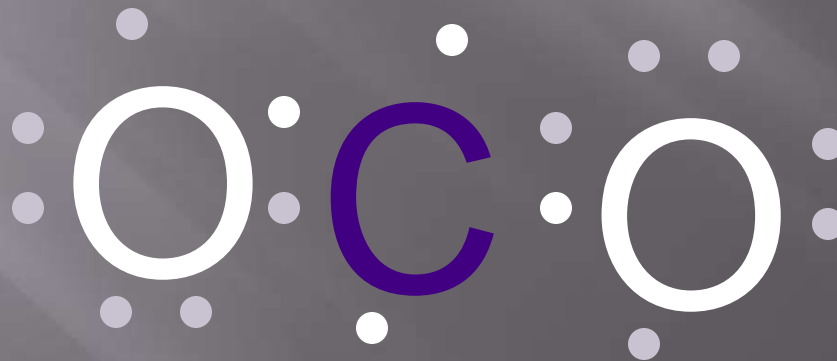
Carbon dioxide

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



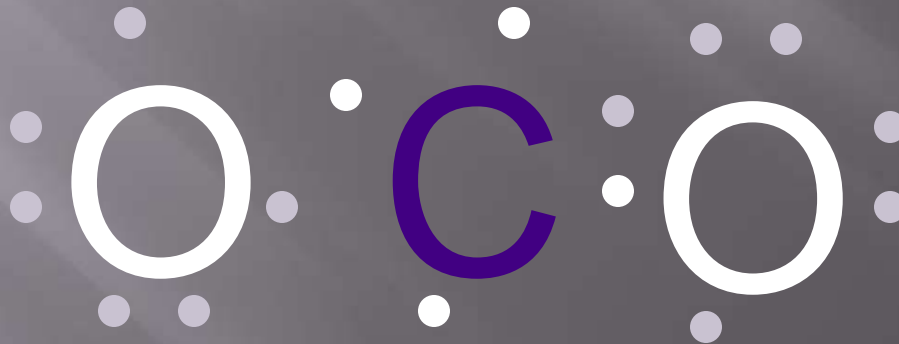
Carbon dioxide

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



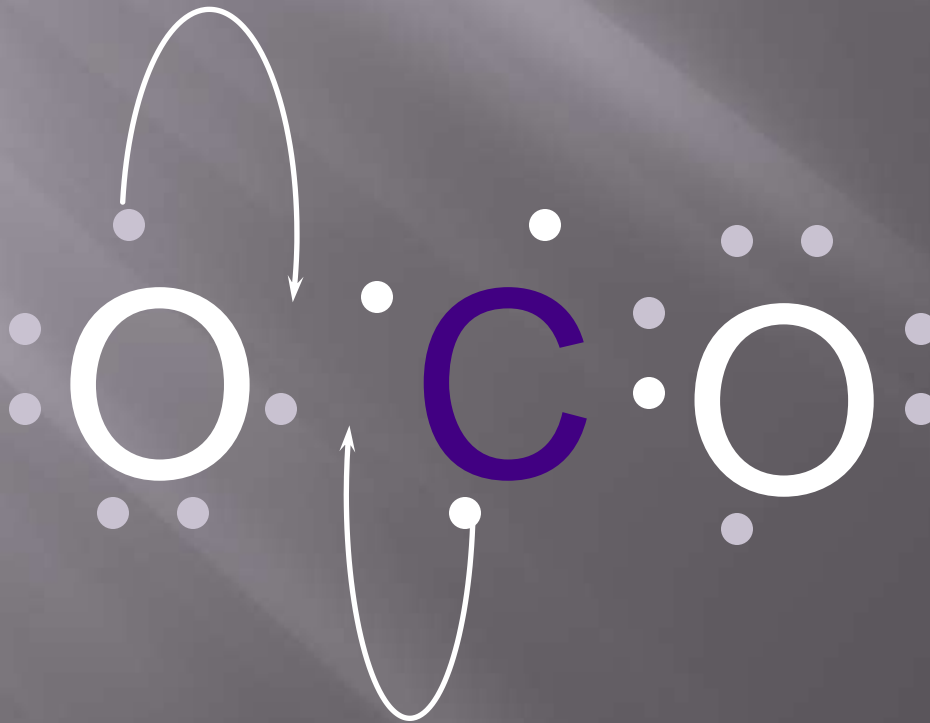
Carbon dioxide

- The only solution is to share more



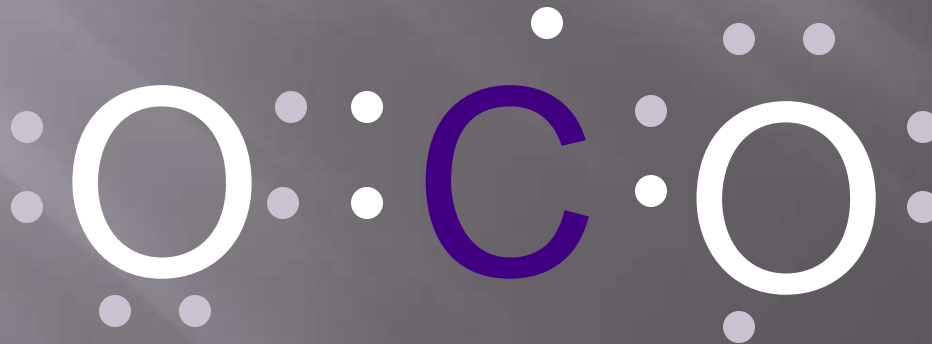
Carbon dioxide

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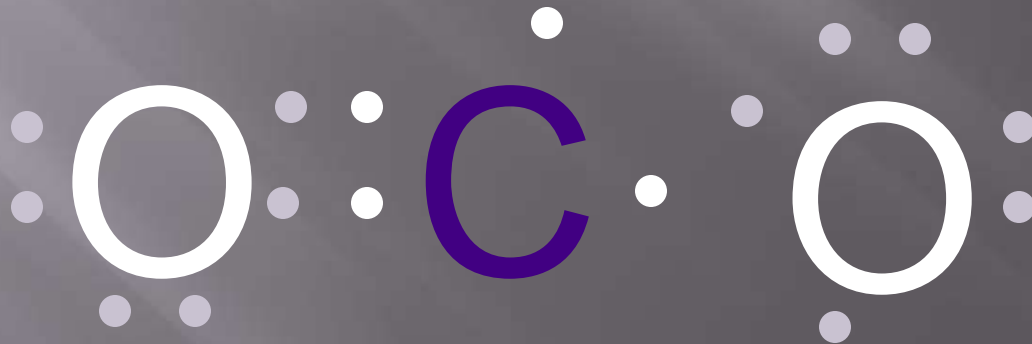
Carbon dioxide

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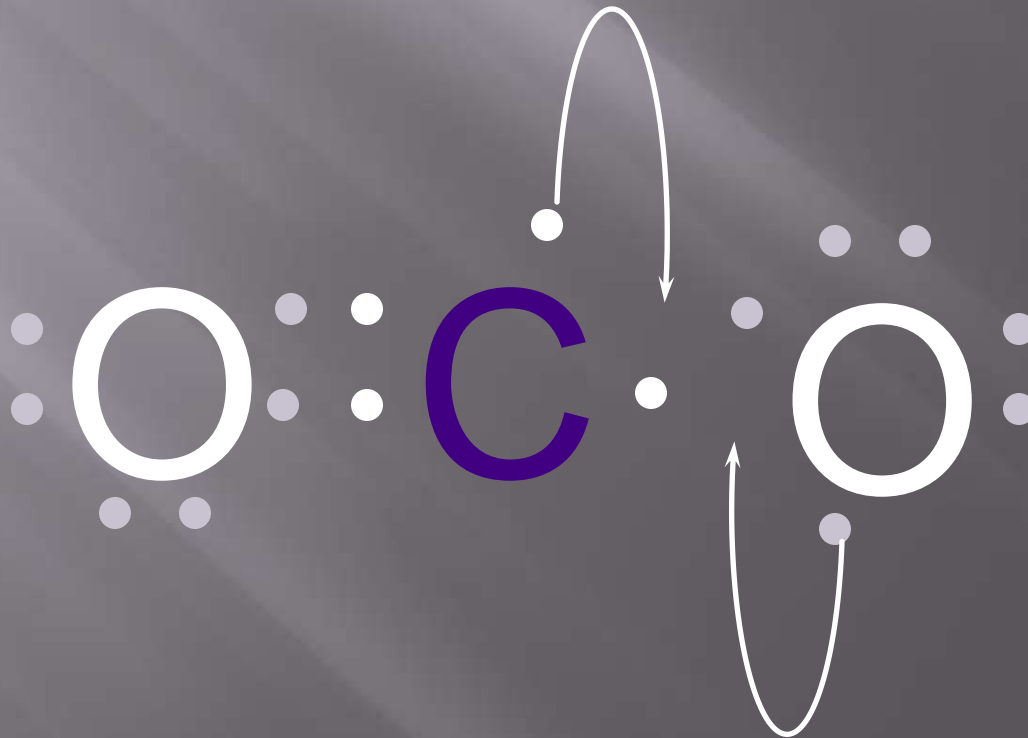
Carbon dioxide

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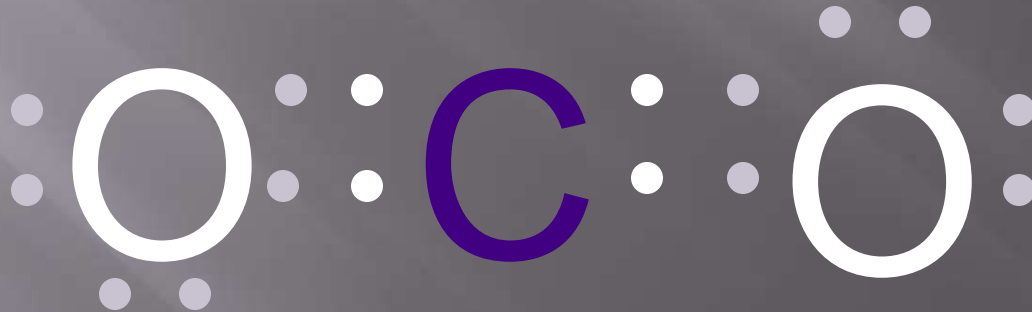
Carbon dioxide

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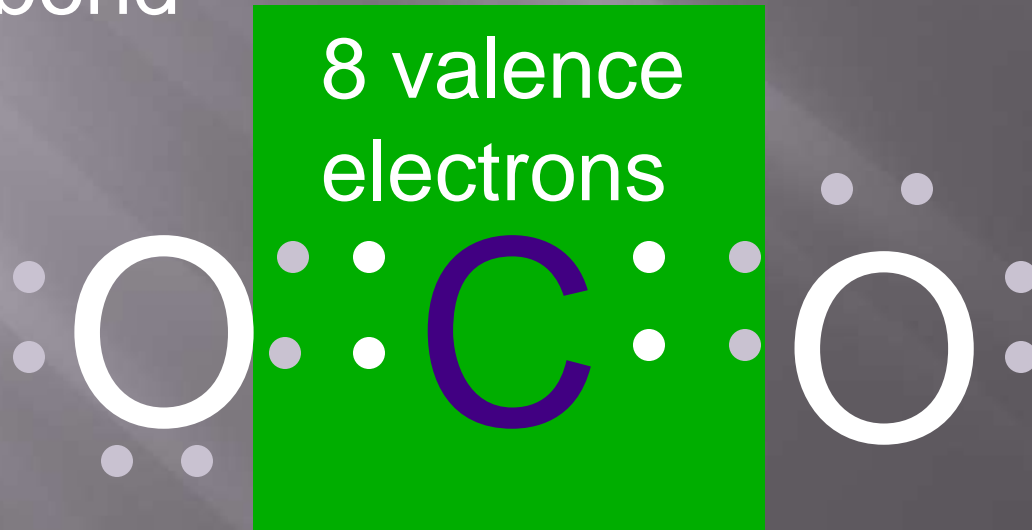
Carbon dioxide

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



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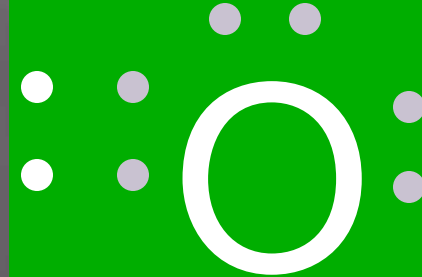


Carbon dioxide

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



8 valence
electrons



How to draw them?

- ▣ Add up all the valence electrons.
- ▣ Count up the total number of electrons to make all atoms happy.
- ▣ Subtract; then Divide by 2
- ▣ Tells you how many bonds - draw them.
- ▣ Fill in the rest of the valence electrons to fill atoms up.

Example

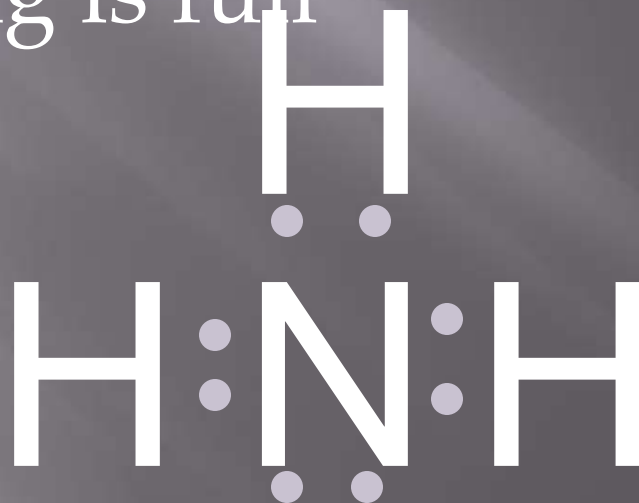
N

H

- NH_3 , which is ammonia
- N - has 5 valence electrons, wants 8
- H - has 1 (x3) valence electron, wants 2 (x3)
- NH_3 has $5+3 = 8$
- NH_3 wants $8+6 = 14$
- $(14-8)/2 = 3$ bonds
- 4 atoms with 3 bonds

Examples

- ▣ Draw in the bonds
- ▣ All 8 electrons are accounted for
- ▣ Everything is full



Example

- ▣ 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 has $5+4+1 = 10$
- ▣ HCN wants $8+8+2 = 18$
- ▣ $(18-10)/2 = 4$ bonds
- ▣ 3 atoms with 4 bonds -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



HCN

- Put in single bonds
- Need 2 more bonds
- Must go between C and N
- Uses 8 electrons - 2 more to add to equal the 10 it has



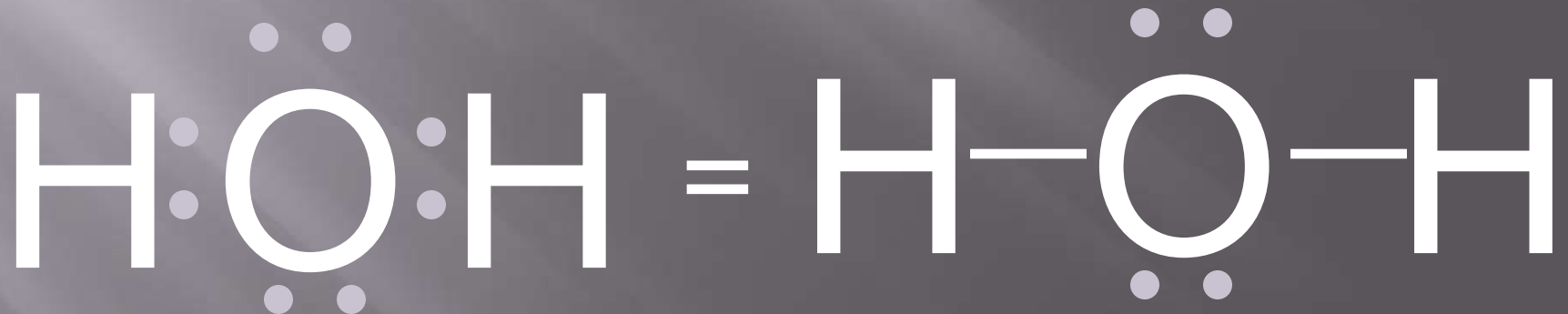
HCN

- 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

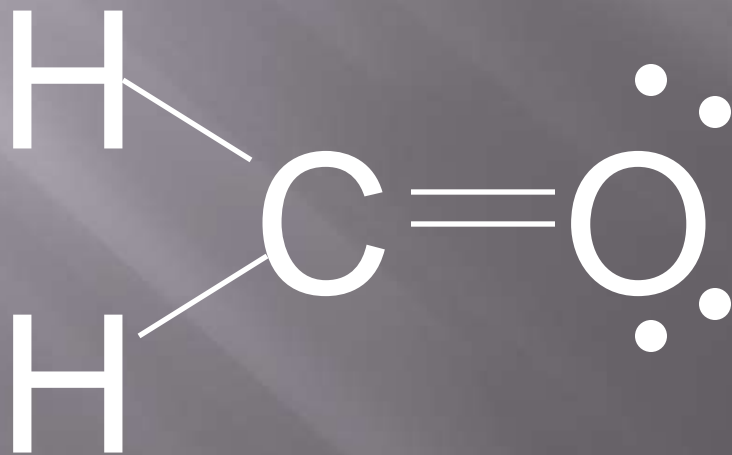
- ▣ Often use a line to indicate a bond
- ▣ Called a structural formula
- ▣ Each line is 2 valence electrons



Structural Examples



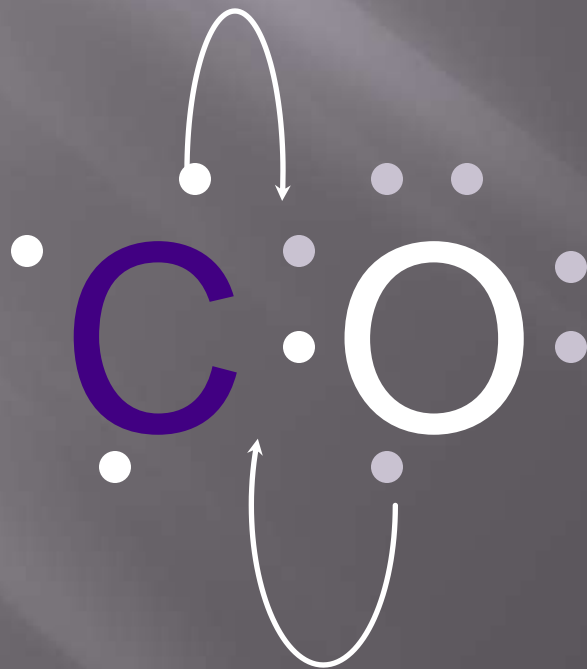
- ▣ C has 8 e⁻ because each line is 2 e⁻
- ▣ same for N



- ▣ same for C here
- ▣ same for O

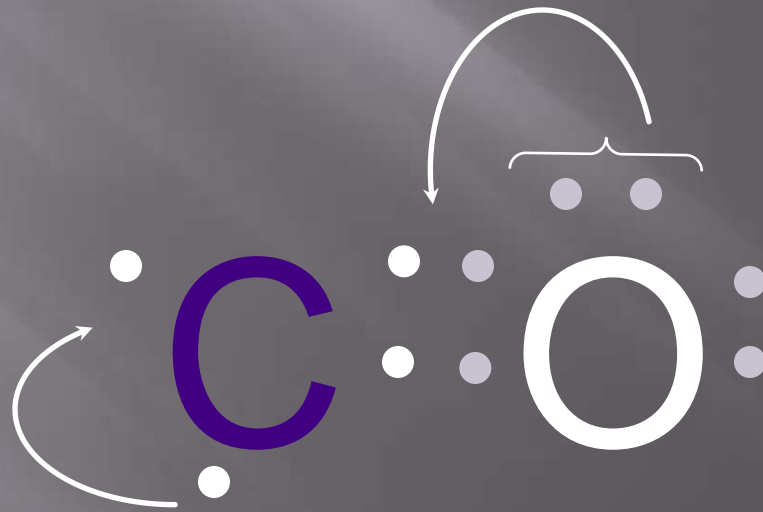
A Coordinate Covalent Bond...

- ▣ 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

Shown as:



Coordinate covalent bond

- ▣ Most polyatomic cations and anions contain covalent and coordinate covalent bonds
- ▣ Add examples here

Bond Dissociation Energies...

- ▣ The total energy required to break the bond between 2 covalently bonded atoms
- ▣ High dissociation energy usually means unreactive
- ▣ Sample: Calculate the kJ to dissociate the bonds in 0.5 mol CO₂

Resonance is...

- ▣ When more than one valid dot diagram is possible.
- ▣ Consider the two ways to draw ozone (O_3)
- ▣ Which one is it?
- ▣ Does it go back and forth?
- ▣ It is a hybrid of both, like a mule; shown by a double-headed arrow

Exceptions to Octet rule

- ▣ For some molecules, it is impossible to satisfy the octet rule
 - usually when there is an odd number of valence electrons
 - NO_2 has 17 valence electrons, because the N has 5, and each O contributes 6
- ▣ impossible to satisfy octet, yet the stable molecule does exist

Exceptions to Octet rule

- ▣ Consider electrons as small, spinning electrical charges
- ▣ creates a magnetic field
- ▣ when paired, they cancel each other, because they are spinning in opposite directions

Exceptions to Octet rule

- ▣ Substances in which all the electrons are paired are called diamagnetic
 - weakly repelled by external magnetic field
- ▣ paramagnetic- substances that contain one or more unpaired e^-
 - attracted to external mag. field

Exceptions to Octet rule

- ▣ Do not confuse with ferromagnetism
 - attraction of Fe, Co, Ni to mag. fld.
- ▣ Oxygen: possible to write structure with all electrons paired
 - not true, because oxygen is paramagnetic
- ▣ Another exception: Boron

Bonding Theories

▣ OBJECTIVES:

- ▣ Describe the molecular orbital theory of covalent bonding, including orbital hybridization.

Bonding Theories

▣ OBJECTIVES:

- Use VSEPR theory to predict the shapes of simple covalently bonded molecules.

Molecular Orbitals are...

- ▣ Orbitals that apply to the overall molecule, due to atomic orbital overlap. 2 types:
 - 1. Bonding orbital - energy is lower than the atomic orbitals from which it is formed
 - 2. Antibonding orbital - energy is higher than what formed them

Molecular Orbitals

- ▣ 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)

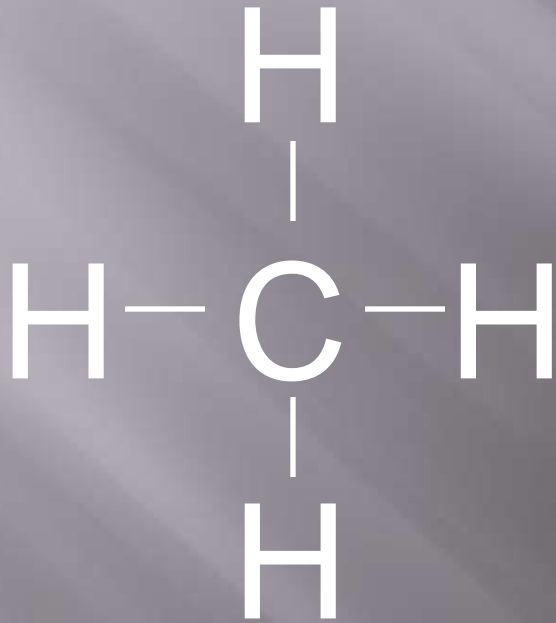
VSEPR: stands for...

- ▣ Valence Shell Electron Pair Repulsion
- ▣ Predicts three dimensional geometry 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.
- ▣ Can determine the angles of bonds.

VSEPR

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

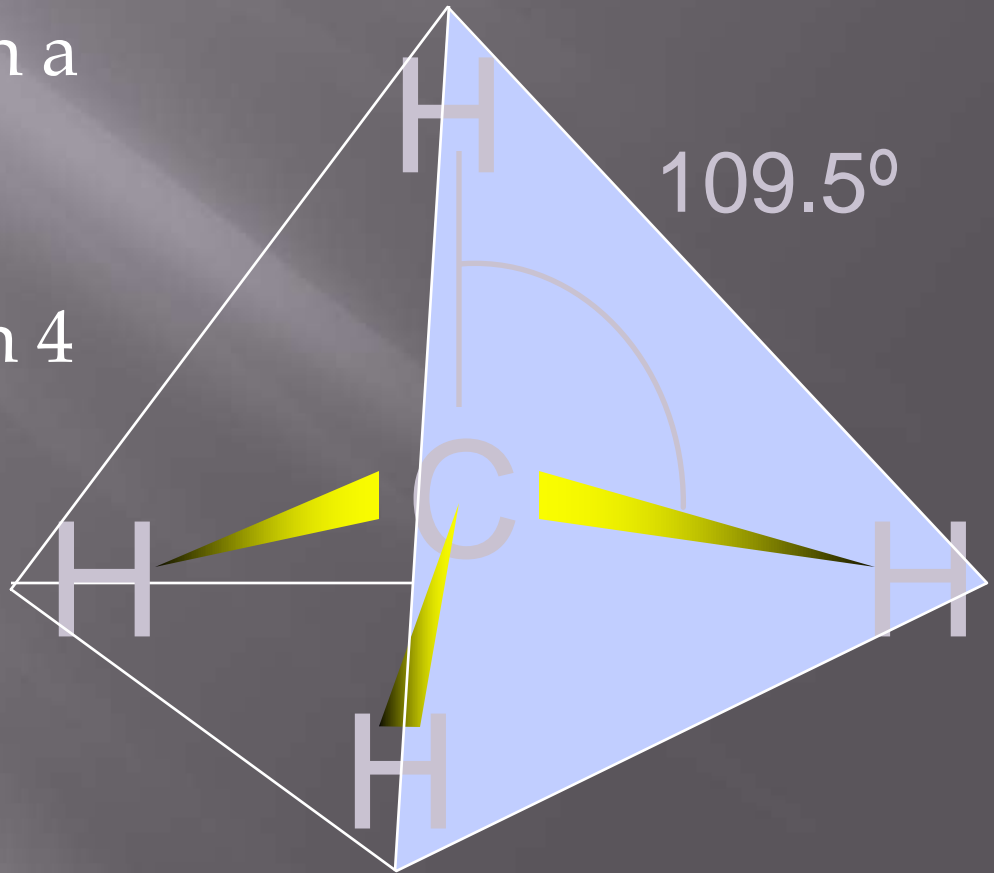
VSEPR



- ▣ 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..

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

- ▣ Note shapes

Polar Bonds and Molecules

▣ OBJECTIVES:

- Use electronegativity values to classify a bond as nonpolar covalent, polar covalent, or ionic.

Polar Bonds and Molecules

▣ OBJECTIVES:

- Name and describe the weak attractive forces that hold groups of molecules together.

Bond Polarity

- ▣ Covalent bonding = 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 different atoms bond covalently, there is an unequal sharing
 - the more electronegative atom will have a stronger attraction, and will acquire a slightly negative charge
 - called a polar covalent bond, or simply polar bond.

Bond Polarity

▣ 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



- ▣ the positive and minus signs (with the lower case delta $\delta^+ \quad \delta^-$) denote partial charges.

Bond Polarity

- Can also be shown:



- the arrow points to the more electronegative atom.

Polarity continued

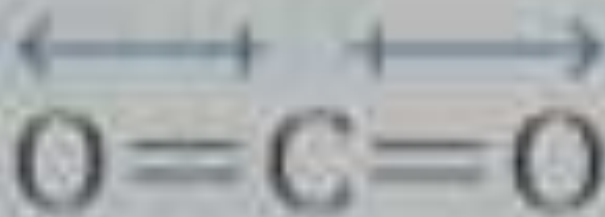
- ▣ If the difference in electronegativity is zero we have a perfectly covalent bond.
- ▣ If the difference in electronegativity between the two atoms in the covalent bond is < 1.7 we have a polar covalent bond
- ▣ If the difference in electronegativity is > 1.7 it “is considered” an ionic bond ???

Polar molecules

- ▣ A polar bond tends 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 dipole

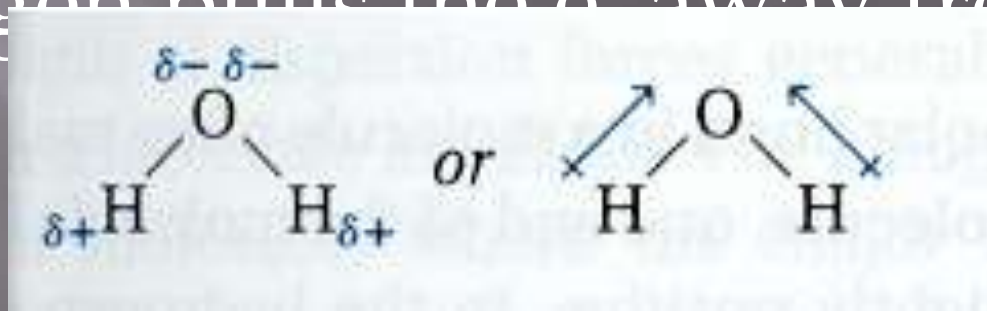
Polar molecules

- ▣ The effect of polar bonds on the polarity of the entire molecule depends on the molecule shape
 - carbon dioxide has two polar bonds, but is linear:



Polar molecules

- ▣ The effect of polar bonds on the polarity of the entire molecule depends on the molecule shape
 - water also has two polar bonds, but the highly electronegative oxygen pulls the e^- away from H:



Attractions between molecules

- ▣ They are what make solid and liquid molecular compounds possible.
- ▣ The weakest 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

2. Dipole interactions

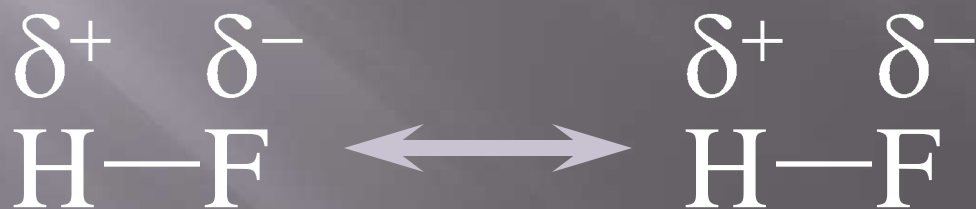
- ▣ Occurs when polar molecules are attracted to each other.
- ▣ Dipole interaction happens in water
 - positive region of one water molecule attracts the negative region of another water molecule.

2. Dipole interactions

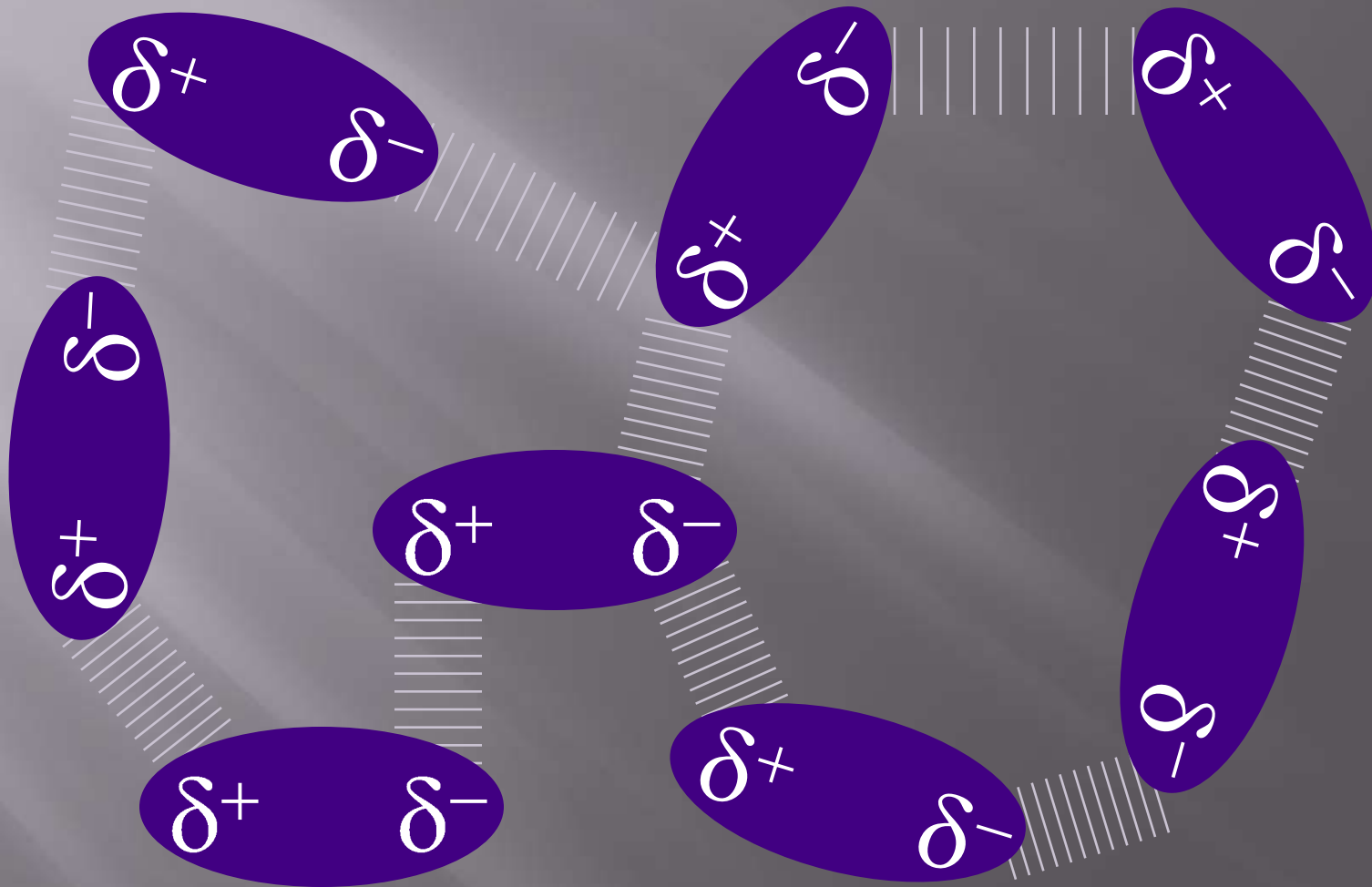
- ▣ Occur when polar molecules are attracted to each other.
- ▣ Slightly stronger than dispersion forces.
- ▣ Opposites attract, but not completely hooked like in ionic solids.

Dipole interactions

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Dipole Interactions



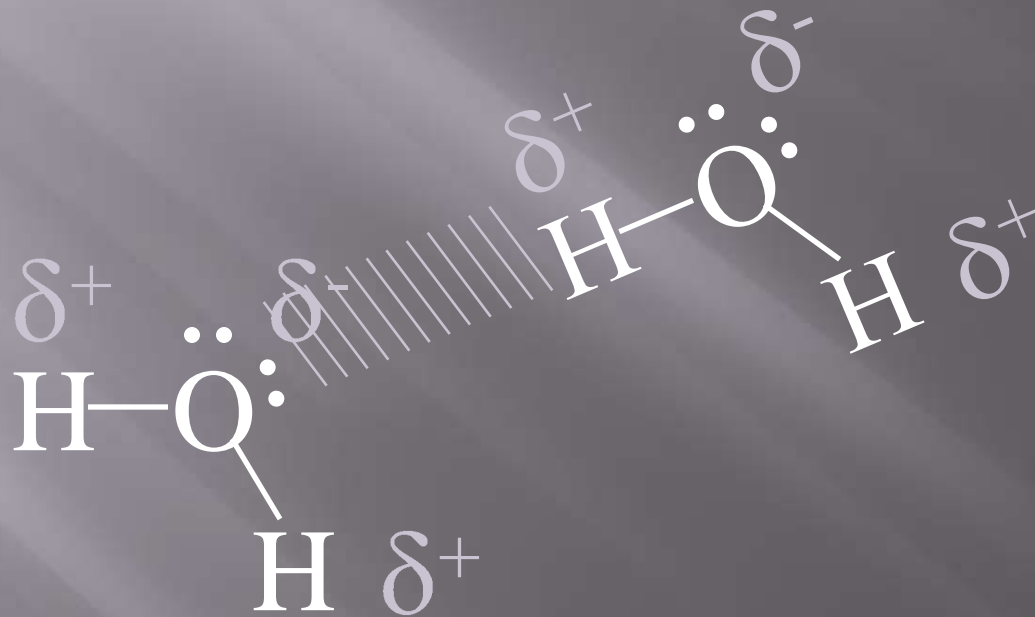
Hydrogen bonding

- ▣ Are the attractive force caused by hydrogen bonded to F, O, or N.
- ▣ F, O, and N are very electronegative so it is a very strong dipole.
- ▣ The hydrogen partially share with the lone pair in the molecule next to it.
- ▣ The strongest of the intermolecular forces.

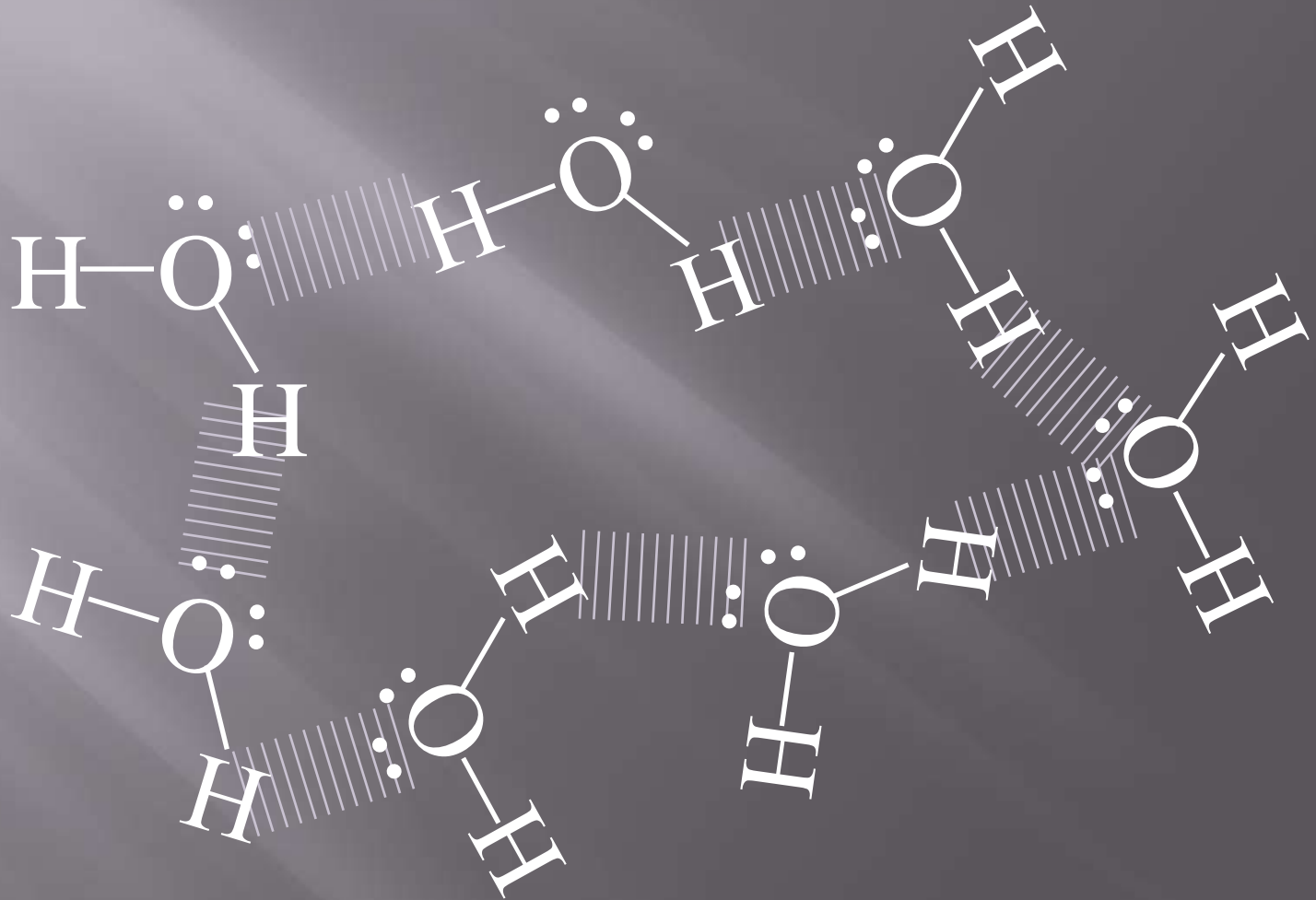
Hydrogen bonding

- ▣ When a hydrogen is covalently bonded to a highly electronegative atom, AND is also weakly bonded to an unshared electron pair of another electronegative atom.
 - The hydrogen is left very electron deficient, thus it shares with something nearby

Hydrogen Bonding



Hydrogen bonding



Attractions and properties

- ▣ Why are some chemicals gases, some liquids, some solids?
 - Depends on the type of bonding
- ▣ Network solids- a special type of molecular solid- melts very high or not at all
 - diamonds, SiC (used in grinding)