# Good Morning

- Grab a whiteboard and draw the diagrams for:
- $CO_2$:
- $H_2O$:
- $SCl_2$:

# Today

- Finish 8.2
- 8.3: Bonding Theories
- Begin 8.4: Polar Bonds and Molecules
- Lab Preview

# Homework

- From Monday: 1-3, 6, 13-15, 20, 21, 39, 41-46, 51.
- From Today (time permitting): 24, 27, 28, 32-35, 37, 38, 54, 57 (58-60)

Homework:

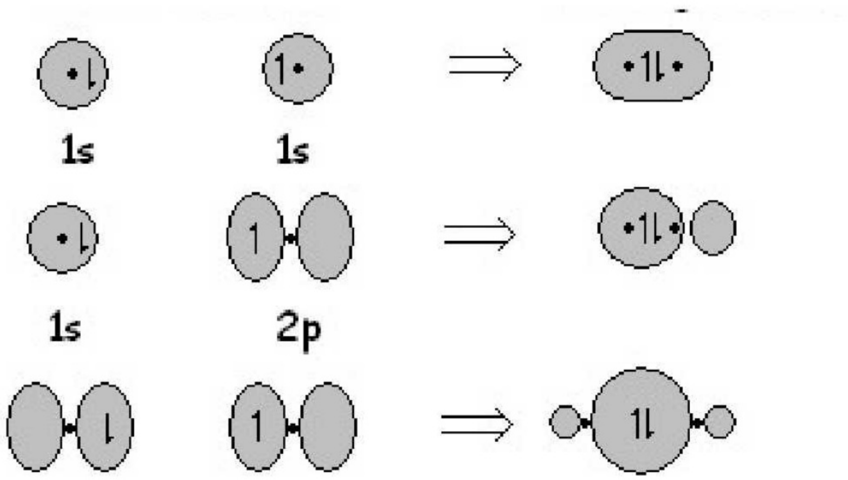
Bonding packet/worksheet.

# 8.3 Bonding Theories

- Relationships between atomic and molecular orbitals.
- The VSEPR theory for predicting the shapes of molecules.
- Orbital Hybridization.

# Molecular Otbitals

- To this point we have assumed that electrons that are shared by two atoms occupy each of the atoms electron orbitals.
- It turns out that the orbitals "merge" to form new shapes.
- These are molecular orbitals.



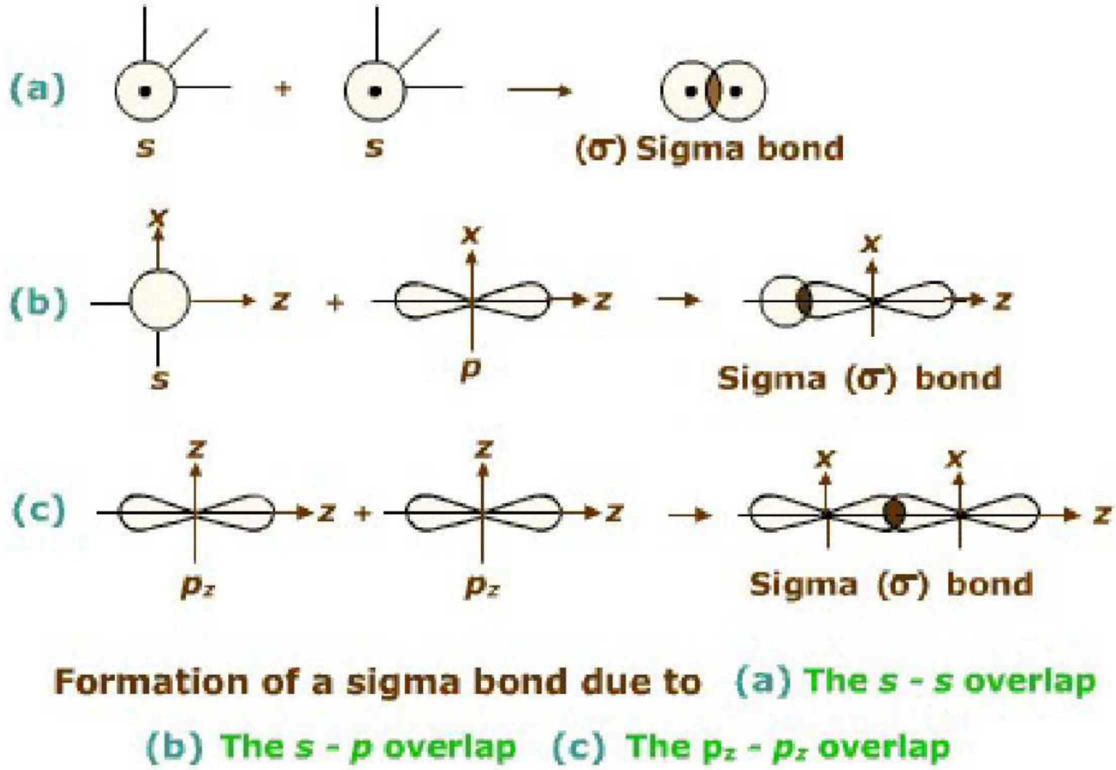

# Examples

# Molecular Orbitals

- These orbitals are shared by the molecule.
- Types:
  - Sigma bond ($\sigma$)
  - Pi bond ($\pi$)

# Sigma Bonds σ

- Two atoms form a **symmetrical** orbital in reference to the "bond axis."
- The bond axis is a line that could be drawn along both nuclei.
- Nuclei are attracted to grater electron density.
- End to end overlap.


(a)
s
s
(σ) Sigma bond
(b)
x
z
s
x
z
p
x
z
Sigma (σ) bond
(c)
z
z
pz
z
z
pz
x
x
z
Sigma (σ) bond
Formation of a sigma bond due to (a) The s - s overlap
(b) The s - p overlap (c) The pz - pz overlap

# Sigma Bonds

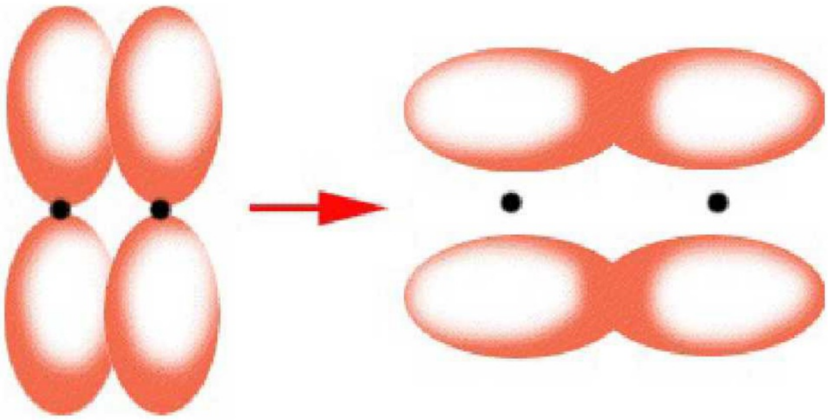
# Examples

- Hydrogen gas:

- Fluorine gas:

# Pi Bonds π

- Side by side overlap.
- Electrons are most likely to be found away from the bond axis.
- "Sausage-shaped"
- Less overlap than sigma bonds.
- Less attraction to nucleus due to proximity.



# Example

# Fact Check

- Pi or Sigma bonds are stronger?
- Why?
- Draw one of each.
- Write questions if you have them.

# VSEPR Theory

- Valence
- Shell
- Electron
- Pair
- Repulsion

# VSEPR Theory

- A better way to represent molecular bonding in 3-D compared to dots.
- Think of Methane $CH_4$
- The hydrogen atoms want to repel each other.

# Formal VSEPR

- The repulsion between electron pairs causes molecular shapes to adjust so that valence electron pairs are as far apart as possible.
- Think like charges repel.

Take a foam sphere.

Make sure that you have a total of 6 push pins.

Put 2 into the spere as far from one another as possible.


VSEPR Theory

# Video Intro

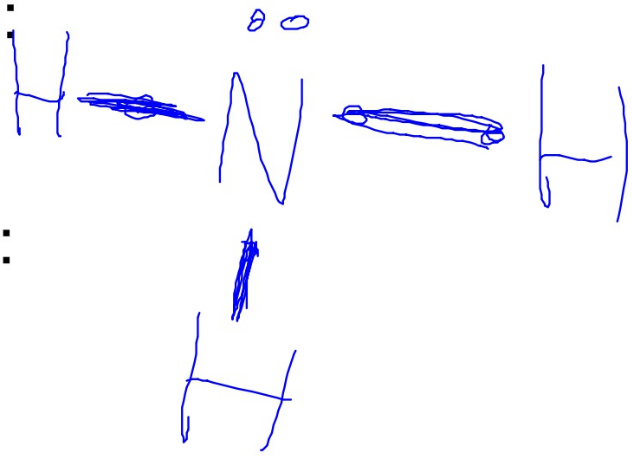
Do Now:

Take out the homework.

In your notes, write the dot structure for

$H_2O$:

8



$NH_3$:

# How Much Space?

- The space between bonding electron pairs depends on many things.
- We will explore the reasons why different atoms make different shapes.
- Depends on the number and types of bonds.

# Shape Shifting

- Shared bonding pairs repel one another.
- Unshared bonding pairs also repel pairs of electrons.
- This will lead to some things being "bent."
- All shapes are listed on p. 233.

# Method

- Draw the dot diagram.
- Note where there are shared electron pairs.
- Note non-shared electron pairs.
- All electron pairs repel.

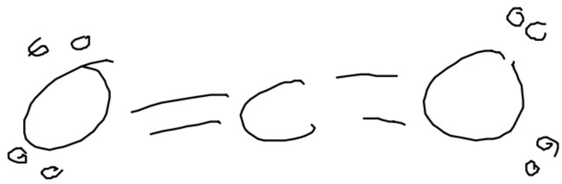
# Triatomic Example

- $H_20$:

2 lone pairs

2 bonded pairs.

- $CO_2$:

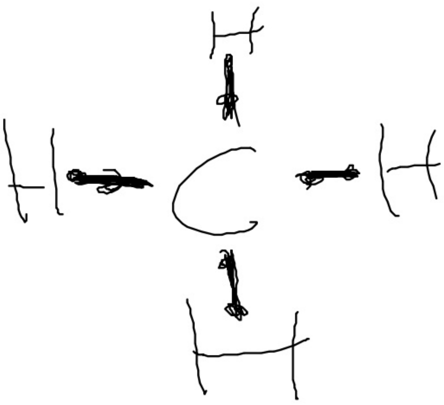


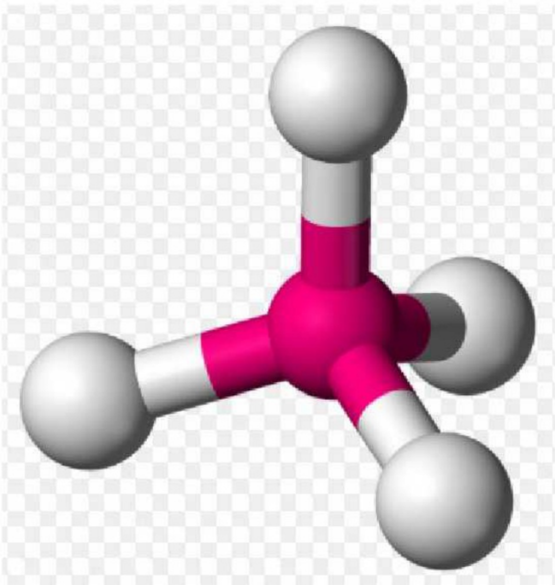
# Tetrahedral

- Tetra-Four
- Hedral-number of flat surfaces on an object.
- These molecules have a central atom and four single covalent bonds.

# Example: Methane

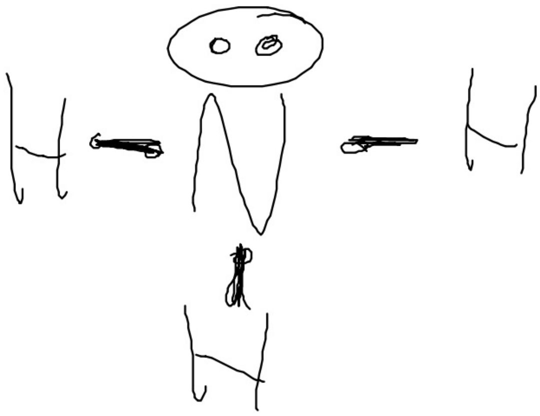
- CH4:

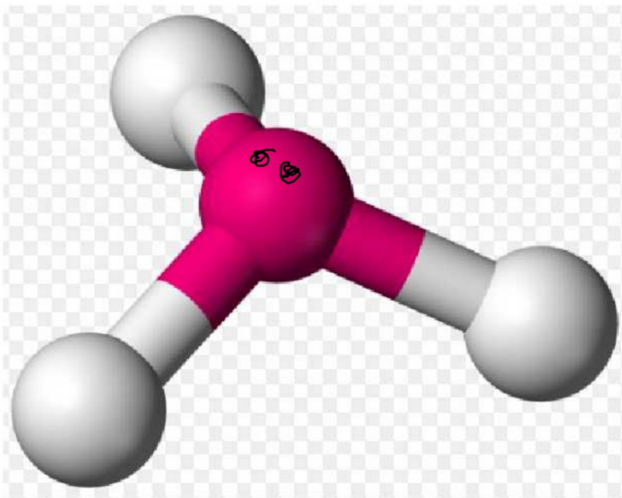


# Pyramidal: Ammonia

- $NH_3$:



Molecules Lab 1:

We will work with the shapes of molecules.

Later, we will fill in the bond polarity and molecule type.

On the back of the page, begin by drawing the Lewis Dot Structures (column one).

Molecule Kits: One per lab table.

Single bonds are definedby a wood dow.

Double and triple bonds require springs.

Spheres are color coated.

C: black H: Yellow O: Red

S: Also Red Cl: Green Br: Orange

N: Light Blue

Build each molecule based on the dot structure that you draw.

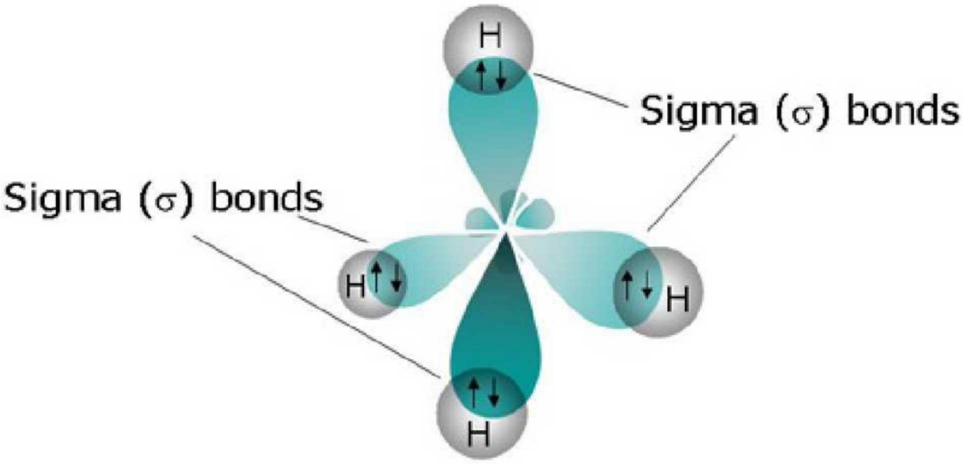
Based on the shape that you build, write the name of the shape in column 2.

# Hybrid Orbitals

- Several atomic orbitals mix together.
- The shapes change.
- They form the same **total** number of orbitals.

# Example: Methane

- Carbon has 4 total valence electrons.
- 2 in the s orbital and 2 in the p orbital.
- Shared s orbitals are stronger than shared p orbitals.
- The s orbital electrons move out to the p orbital and the shape is symmetrical.




# Methane

# 8.4 Polar Bonds and Molecules

- How does electronegativity determine the distribution of charge in a molecule?
- What happens when polar molecules are between metal plates?

# 8.4 Polar Bonds and Molecules

- Strength of intermolecular attractions in molecules compared to ionic compounds.
- Why do "network solids" have a high melting point?


Polar & Non-Polar Molecules: Crash Course Chemistry #23

# Bond Polarity-4:45

# Non-polar Covalent

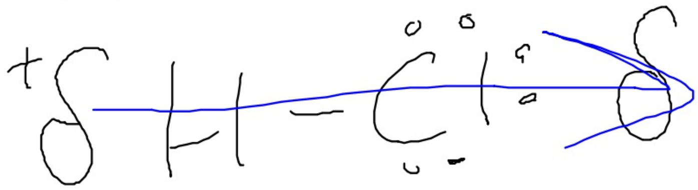
- The pull of each atom on the electrons that it shares is equally distributed.
- These have a balance of charge.
- Diatomic Molecules like hydrogen and oxygen gas are examples of this.

# Polar Covalent Bonds

- One atom has a higher electronegativity than another atom.
- The more electronegative atom attracts the electrons more.
- One region is more negative than another.

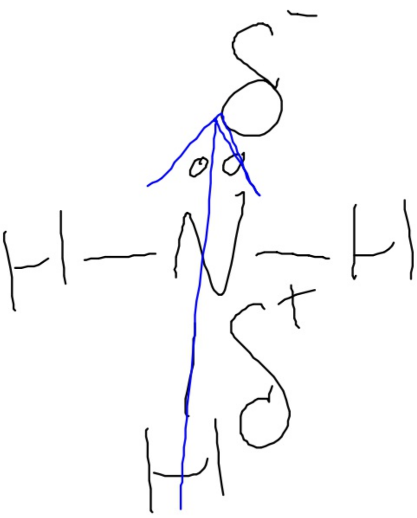
Labeling polar molecules:

HCl:



H2O:

NH3:



# Difference in Electronegativity

- 0.0-0.4: Non-polar covalent bond.
- 0.4-1.0: Moderately polar covalent.
- 1.0-2.0: Very polar covalent.
- More than 2.0: Ionic bond.

| | | | | | | | | | | | | | | | | | |
|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|
| 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 1.8 | 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 3.0 |
| 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 2.6 |
| Cs 0.7 | Ba 0.9 | La 1.1 | 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 1.8 | Bi 1.9 | Po 2.0 | At 2.2 | Rn 2.4 |
| Fr 0.7 | Ra 0.7 | Ac 1.1 | Unq | Unp | Unh | Uns | Uno | Une | | | | | | | | | |

| | | | | | | | | | | | | | |
|---|---|---|---|---|---|---|---|---|---|---|---|---|---|
| Ce 1.1 | Pr 1.1 | Nd 1.1 | Pm 1.1 | Sm 1.1 | Eu 1.1 | Gd 1.1 | Tb 1.1 | Dy 1.1 | Ho 1.1 | Er 1.1 | Tm 1.1 | Yb 1.1 | Lu 1.2 |
| Th 1.3 | Pa 1.5 | U 1.7 | Np 1.3 | Pu 1.3 | Am 1.3 | Cm 1.3 | Bk 1.3 | Cf 1.3 | Es 1.3 | Fm 1.3 | Md 1.3 | No 1.3 | Lr |

# PT Electronegativity

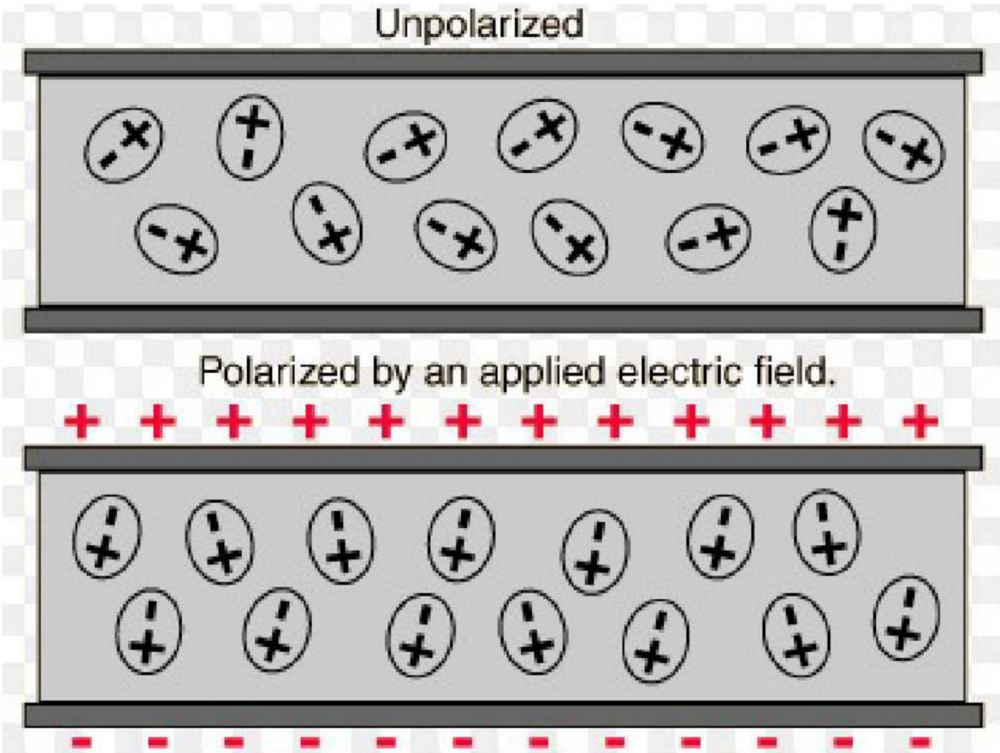
# Water

Do Now:

Take out the homework from last night.

Please take out the PTE with electronegativity, your notebook, and something to write with.

Answer the following in your notes:  Based on what you know about ionic bonding, what affect does the polarity of a molecule have on the way that it interacts with other compounds.




# Polar Molecules Between Charged Plates

Intermolecular forces

These are "mild" forces between molecules that make some substances want to stick together a little bit.

Finish Crash Course 23

# Dipole Effect

- Polar molecules have charges on different sides.
- Dipole: two poles.
- These things respond to electrical charges.

# Attractions Between Molecules

- Van der Waals Forces - The two weakest forces between groups of the same kind of molecule.
- 1) Dipole interactions - positive end of one molecule is attracted to the negative end of another molecule of the same kind.

# Attractions Between Molecules

- 2) Dispersion Forces - electrons from one molecule affect the electrons of another molecule.
- Example: + + + + +

# Hydrogen Bonds

- Hydrogen that is bonded in a molecule tends to have it's 1 electron towards the center of the molecule.
- The part facing out is more positive.
- This is attracted to negative parts of molecules.

# Intermolecular Attractions

- These are "network solids"
- All of the atoms are covalently bonded to each other.
- These molecules bond to each other and for lattice structures the way that ionic compounds do, but with covalent bonds instead of electrostatic forces.

# Physical Properties

- VERY high melting point.
- Not water soluble.
- Very low electrical conductivity.
- Ex: Diamond do not melt. They vaporize at 3500 degrees Celsius.

Today's Lab

Building Molecules II.

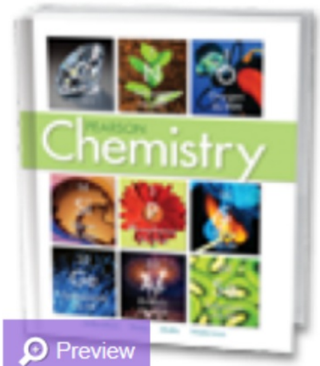
Get a molecule building kit for each table.

Answer both sides of the worksheet and put in in your lab folder when you are finished.

Finished?

Take out the table that compares metallic, ionic and ocavlent bonding.

Complete the table.




Preview

# Duplicate of Covalent Bonding

Chemistry Covalent Bonds

Play ▸ Preview Favourite ★

Share

Or, copy & share this link: https://play.kahoot.it/#/k/8dbd4fdf-6910-41c

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