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VCE Chemistry ½
Covalent Molecules [1.6]
Workbook

Outline:

Covalent Bonds Pg 2-26
- Lewis Structures
- Covalent Bonds
- Drawing Lewis Structures of Molecules
- Deriving Molecular Formula
- Multiple Covalent Bonds

Shapes of Molecules Pg 27-46
- Molecular Geometry
- Molecular Geometry with Double and Triple Covalent Bonds
- Parent/Electron Geometry vs Molecular Geometry

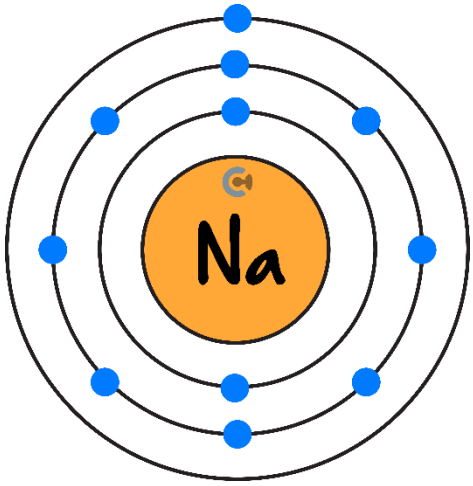
Learning Objectives:

- ☐ **CH12 [1.6.1]** - Draw Lewis structures of atoms & covalent molecules.
- ☐ **CH12 [1.6.2]** - Identify the geometries (parent & molecular) of molecules, with reference to VSEPR theory.

Section A: Covalent Bonds

Sub-Section: Lewis Structures

Discussion: What might be an issue with drawing shell model diagrams?

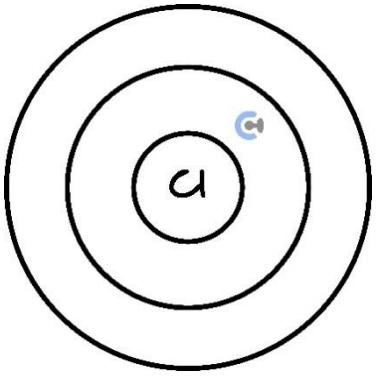


Exploration: Chlorine Shell Model Diagram

$^{35}_{17}Cl$

- How many protons and electrons does chlorine have?
- _____ protons and _____ electrons.
- What is the Shell Model electron configuration of chlorine?
- How many **valence** electrons does chlorine have?

- What is the shell model diagram for chlorine? *(Label Below)*



NOTE: We don't care about inner shell electrons. The number of _____ electrons matter a lot more.

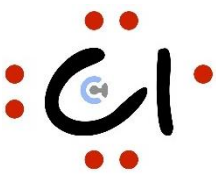
Exploration: Chlorine Lewis Structure

- The following short-formed version of the shell model diagram is used instead *(Label Below):*

- This is called the _____ of an atom, and it only shows the number of _____ electrons in an atom.

Lewis Structure / Lewis Dot Structure

- **Definition:** The Lewis Structure only shows the **valence electrons** in an atom, which is typically shown by _____.
- For example:



NOTE: Lewis structures are used because it's much **quicker** to draw molecules compared to the Shell Model Diagram.

Let's practise this together!

Question 1 Walkthrough.

Draw the Lewis Structure for Lithium (Li).

Your Turn!

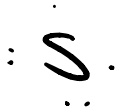
Question 2

Draw the Lewis Structures for each of the following atoms:

a. Nitrogen (N).

b. Carbon (C).

c. Sulphur (S).



Question 3 (1 mark) **Additional Question.**

Draw the Lewis dot structure for bromine.

Misconception

" The Lewis Structure of atoms can be drawn in any way."

- Sulphur can be written in any of these two ways:

TRUTH: Lewis Structures can' t be drawn in any way.

Misconception

" Electrons always pair up first."

TRUTH: Electrons do pair up, but they pair up last, after the 4 valence orbitals are half-filled first.

What does this look like in practice?

Question 4 Walkthrough.

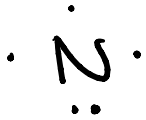
Draw the Lewis Structures for Nitrogen (N).

TIP: Start by drawing electrons on all **four sides** (top, bottom, right, left) of the atom before starting to pair the electrons up.

Definition

Electron Pairing in Lewis Dot Structures

- Electrons pair up last after the 4 valence orbitals are **half-filled** first.
- Nitrogen's Lewis Structure:



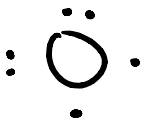
NOTE: The electrons in Lewis Structures can be drawn as either **dots or crosses.**

Try some yourself!

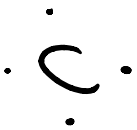
Question 5

Draw the Lewis Structures for each of the following atoms:

a. Oxygen (O).

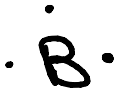


c. Carbon (C).



b. Phosphorus (P).

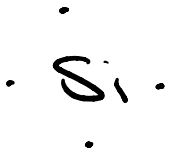
d. Boron (B).



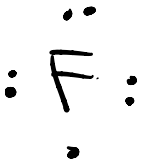
Question 6 Additional Question.

Draw the Lewis Structures for each of the following atoms:

a. Si



b. F



Sub-Section: Covalent Bonds

Active Recall: According to the Octet Rule, how many electrons do atoms want in their outer shell?

8e⁻ in our shell → stable

Exploration: Chlorine Obtaining a Full Outer Shell

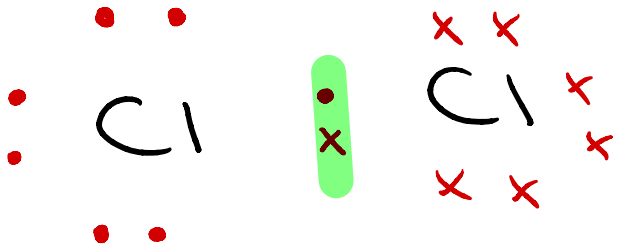
- Chlorine has 7 valence electrons.
- What will a chlorine atom try to do to achieve a full outer shell?

  [Gain] / [Lose] an electron.
- What will this look like on a chlorine atom? *(Label Below)*

- If there are only other chlorine atoms nearby, what is the issue with each of them? *(Label Below)*

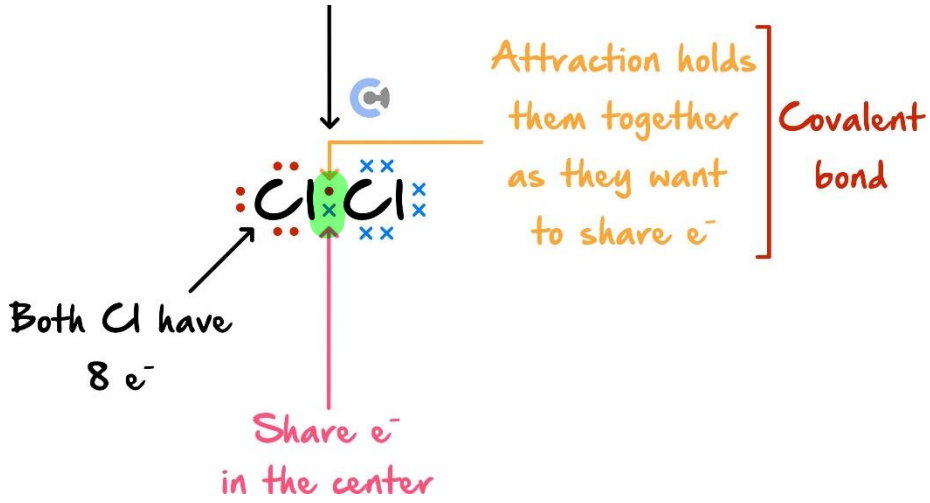
- As such, will either atom want to give up an electron?
  [Yes] / [No]
- So, what do they end up doing? *(Label Below)*
- How many electrons do both chlorine atoms have now? *(Label Below)*

- What is this bond called? *(Label Below)*



Covalent Bond

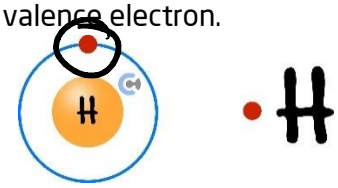
- **Definition:**
  - A chemical bond where two or more atoms ~~share~~ _____ electrons.
  - The covalent bond is formed so that all atoms have **full outer shells**.



Let's explore this idea with another element!

Exploration: Hydrogen Atom Covalent Bonding

- Consider hydrogen (H) which has 1 valence electron.

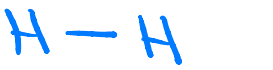


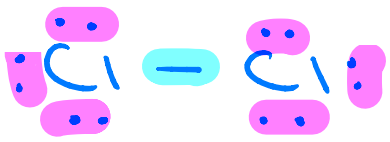
- How many electrons will make its valence electron shell full? 2
- As such, how many electrons does it need to obtain a full outer shell? 1
- As such, what happens if two hydrogen atoms are placed close to one another? *(Label Below)*

NOTE: A simplified way to represent a covalent bond is to use a _____ to connect the two atoms involved in the bond.

Exploration: Molecules

- As hydrogen and chlorine need one more electron, they can share one electron with each other:





- The electrons in atoms involved in bonding can be classified into two categories:
  - bonding electrons.
  - non-bonding electrons.
- What does this look like on chlorine and hydrogen? *(Label Above)*
- The name given to the final 'group of atoms' is a molecule.

Bonding and Non-bonding Electrons

- Bonding electrons are valence electrons which are **directly** involved in a covalent bond.
- Non-bonding electrons are electrons which are **not directly** involved in a covalent bond.

NOTE: Covalent bonding always occurs between non-metals, as they need to **gain** electrons!

Discussion: What is a molecular formula?

Exploration: Molecular Formula

- The molecular formula lists the number of each atom in a molecule.
- What is the molecular formula of each of the following? *(Label Below)*:

H_2

HCl

Cl_2

Molecule

- **Definition:**
  - A group of **two or more** atoms held together by **covalent** bonds.
- The molecular formula lists the **amount** of each atom in a **molecule.**

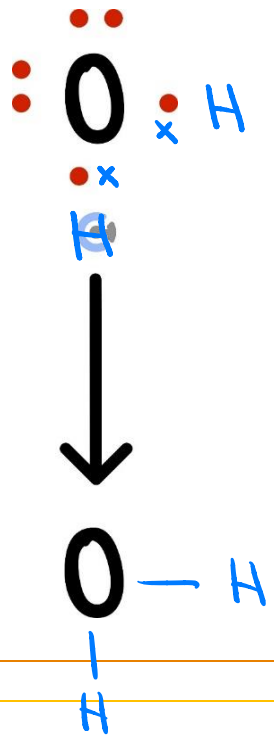
What if the molecular formula is given and the Lewis Structure needs to be derived from it?

Exploration: Water

- What's the molecular formula for water?
  - H_2O
- Lewis Structure for oxygen (O): *(Label Below)*

- How many more electrons does oxygen need to obtain a full outer shell?
  - 2
- How many covalent bonds does oxygen have to form to obtain a full outer shell?
  - 2

- The oxygen atom will bond with the two hydrogen atoms as follows: *(Label Below)*



NOTE: The number of covalent bonds each atom forms is equal to the number of electrons it requires to obtain a full outer shell.

- One covalent bond = One more electron gained by sharing

ALSO NOTE: Within a molecule, there may be **multiple** covalent bonds.

Let's have a look at a question together!

Question 7 Walkthrough.

For nitrogen (N), indicate the maximum number of covalent bonds which will form.

3

Space for Personal Notes

Your Turn!

Question 8

For each of the following atoms, indicate the maximum number of covalent bonds which will form:

a. Oxygen (O).

2

b. Carbon (C).

4

c. Hydrogen (H).

1

d. Fluorine (F).

1

e. Neon (Ne).

0

f. Argon (Ar).

0

NOTE: Noble gases such as Ne or Ar have a full outer shell, thus do not form bonds with anything and are unreactive as a result!

Sub-Section: Drawing Lewis Structures of Molecules

Let's put what we've learnt so far together in a question!

Question 9 Walkthrough.

Draw the Lewis Structure for ammonia (NH_3).

N → 5 → 3
H → 1 → 1

TIP: When trying to draw out a molecule, draw the Lewis structure first, and figure out how many covalent bonds each atom will try to form!

Your Turn!

Question 10

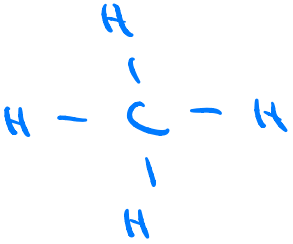
Draw the Lewis Structure for each of the following molecules:

a. Hydrogen sulphide (H_2S).

S → 2
H → 1

b. Methane (CH_4).

C → 4
H → 1

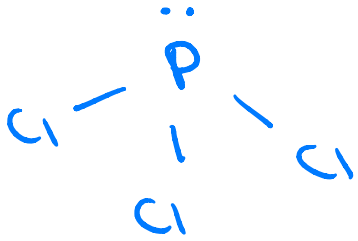


Question 11

Draw the Lewis Structure for each of the following molecules:

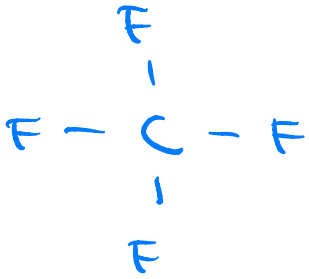
a. Phosphorus trichoride (PCl_3).

P → 3
Cl → 1



b. Carbon tetrafluoride (CF_4).

C → 4
F → 1



Question 12 Additional Question.

Draw the Lewis Structure for each of the following molecules:

a. Methane (CH_4).

b. Ethane (C_2H_6).

Sub-Section: Deriving Molecular Formula

What if we need to work the other way?

Question 13 Walkthrough.

Identify the most likely formula of the molecule formed between Carbon (C) & Chlorine (Cl).

C → 4 Cl → 1

CCl_4

TIP: Figure out how many covalent bonds each atom will try to form and find the simplest combination that gives each atom a full outer shell!

Your Turn!

Question 14

Identify the most likely formula of the molecule formed between the following pairs of elements:

a. Hydrogen & Iodine.

H → 1

I → 1

HI

b. Nitrogen & Bromine.

N → 3

Br → 1

NBr_3

Question 15 Additional Question.

Identify the most likely formula of the molecule formed between the following pairs of elements:

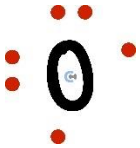
a. Sulphur (S) & Fluorine (F).	**b.** Silicon and Hydrogen.

Sub-Section: Multiple Covalent Bonds

Discussion: Can atoms share multiple pairs of electrons?

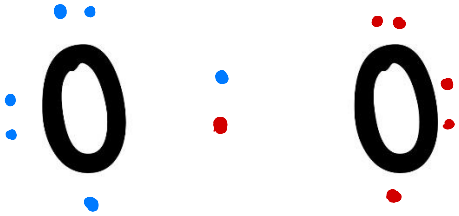
Exploration: Multiple Covalent Bonds

- Consider oxygen again:



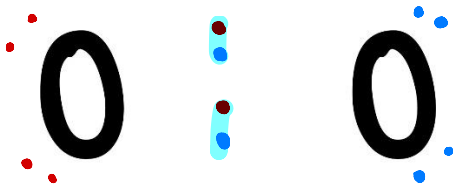
Valence Electrons	Electrons Required	Covalent Bonds Formed
6	2	2

- What if it were to form a covalent bond with itself? *(Label Below)*



- This is a single covalent bond, is this the ideal formation? [Yes] / [No]

- What if the atoms shared **two** pairs of electrons? *(Label Below)*



- This covalent bond is a [Single] / [Double] bond, is this the ideal formation? [Yes] / [No]

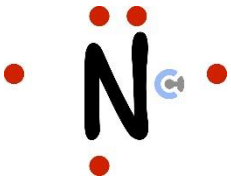
Multiple Covalent Bonds

- Double covalent bonds are the same as a pair of atoms bonding with each other twice.
- Can also be denoted by a **double straight line**:

What about more covalent bonds?

Exploration More Covalent Bonds

- Consider nitrogen:

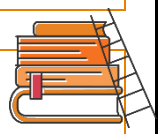


- How many more bonds does it want to form?

3

- Nitrogen (N) atoms can bond with itself. *(Label Below):*

- triple covalent bonds can also form!



Extension: Quadruple covalent bonds

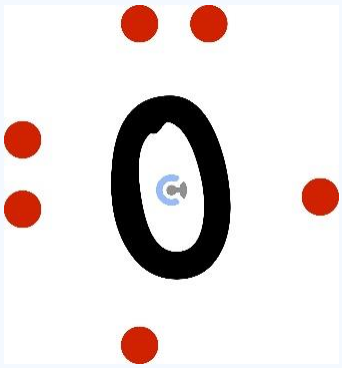
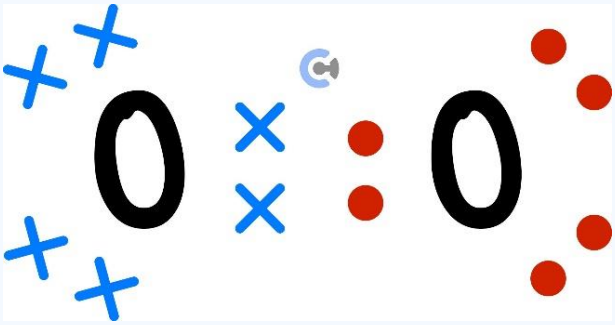
- They exist but are not covered in VCE Chemistry!

Discussion: **In nature, are elements more likely to exist as singular atoms, or are they more likely to exist in pairs?**

Let's take a look at this idea!

Exploration: Elements in Nature

- Consider a single oxygen atom compared to an oxygen molecule (O_2).

Oxygen Atom (O)	Oxygen Molecule (O_2)
	

- The oxygen [atom] / [molecule] contains full outer shells.
- Hence, the oxygen [atom] / [molecule] is more likely to exist in nature.

NOTE: Molecules are more likely to exist in pairs such as O_2.

ALSO NOTE: This is because atoms would rather pair up and form a diatomic molecule where the atoms have full outer shells.

Diatomic molecule

- **Definition:**
  - A molecule with exactly two atoms in its structure.
  - These molecules exist naturally as they have full outer shells.

TIP: The following mnemonic helps remember which elements are diatomic in nature:

"Have No Fear Of Ice Cold Beer."

Let's explore this further!

Exploration: Carbon Dioxide (CO_2) Structure

- In molecules such as carbon dioxide (CO_2), there could be single, double or triple bonds!
- Let's look at carbon dioxide:

Element	Carbon (C)	Oxygen (O)
Electron Configuration	2,4	2,6
Number of electrons required to obtain a full outer shell	4	2
Number of covalent bonds each atom can form	4	2

- Carbon can form more bonds, it is likely to be in the middle of the molecule.
- What will the structure of CO_2 look like? *(Label Below)*

- What are the shared and other electrons called? *(Label Below)*

TIP: Figure out what requires the most electrons – that is the atom(s) which are most likely to be in the middle of the molecule!

Let's have a look at a question together!

Question 16 Walkthrough.

Draw the Lewis Structure for Hydrogen Cyanide (HCN).

H → 1
C → 4
N → 3

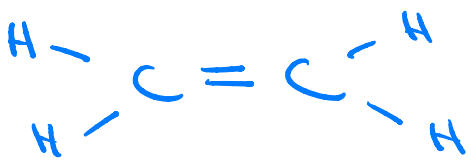
Your Turn!

Question 17

Draw the Lewis Structure for:

a. Ethene (C_2H_4).

C → 4
H → 1



b. Ethyne (C_2H_2).

Question 18 Additional Question.

Draw the Lewis Structure for nitroxyl (HNO).

H → 1
N → 3
O → 2

Space for Personal Notes

Key Takeaways

- The Lewis Structure shows **valence electrons** in an atom, typically depicted by **dots/crosses**.
- A **covalent** bond is a chemical bond where two or more atoms **share** electrons.
- **Bonding** electrons are valence electrons which are directly involved in a covalent bond.
- **Non-bonding** electrons are electrons which are not directly involved in a covalent bond.
- Covalent bonding always occurs between **non-metals**, as they need to gain electrons!
- The **molecular** formula essentially lists the **amount** of each atom in a molecule.
- In nature, many elements exist in their **diatomic** form, as the atoms have **full** outer shells.
- The element that can form the **most** covalent bonds is usually in the **middle** of the molecule.

Space for Personal Notes

Section B: Shapes of Molecules

Sub-Section: Molecular Geometry

Context

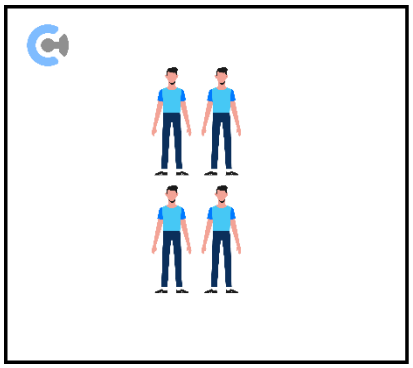
- While we've been looking at the Lewis structures along with the bonding structure of molecules, we have not looked at the **shape** of these molecules!

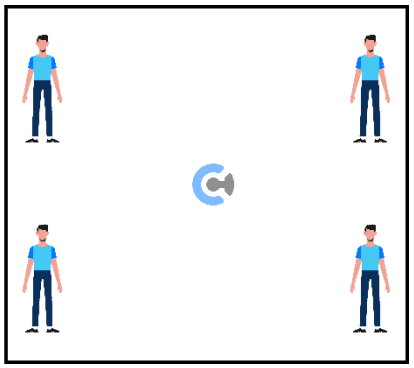
Molecular Geometry

- The 3D arrangement of atoms that constitute a molecule.

Analogy

- Imagine four people **who hate each other** have been forced into a room together.
- How are they most likely to position themselves in the room? *(Circle Below)*





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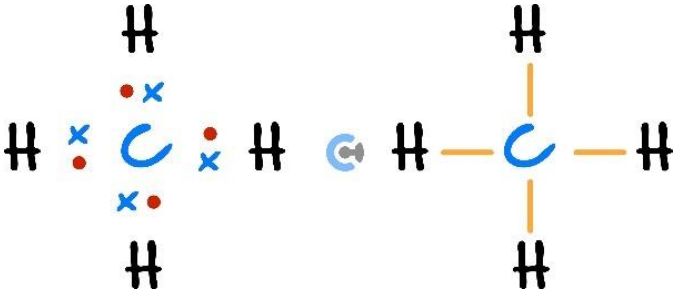
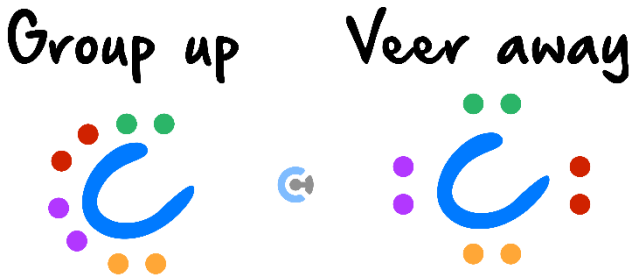
Let's take a look at this idea with molecules!

Exploration: Shape of Methane (CH_4)



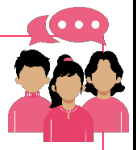
- Consider the four electron pairs/orbitals surrounding carbon. *(Label Below)*

- The charge on each of these electron pairs is [positive] / [negative].
- The four electron pairs will [attract] / [repel] each other.
- Are the electron pairs more likely to group up or separate as much as they can? *(Circle Below)*

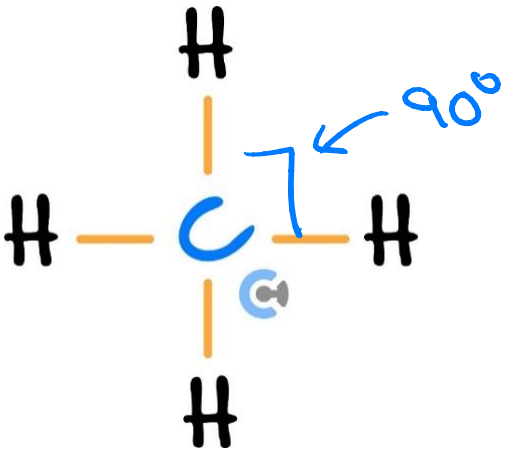


Electron Arrangement

- As electron pairs repel each other, they are located as ____as possible from each other.

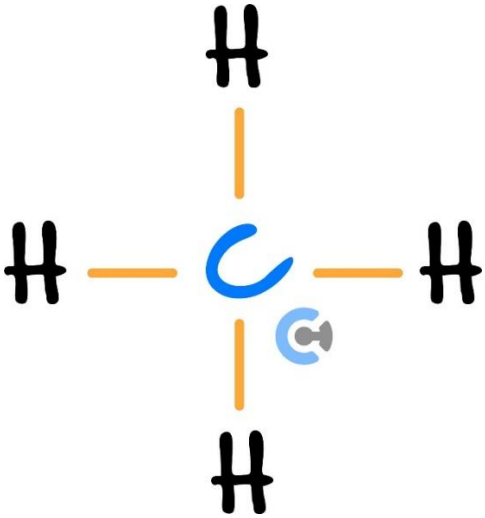


Discussion: What angle will exist between each of the bonds in methane (CH_4)?



Misconception

" When the electron pairs separate from each other as much as possible, they will form a cross formation as shown, with bonds located 90° away from each other.



TRUTH: This is not the formation which results in bonds being located the farthest away from each other!

- This is true for a 2D plane, but as atoms exist in 3D, there are other formations which result in electron groups being located **even further away** from each other!

Let's have a look at some molecules' shapes!

Exploration: Shapes of Molecules

- **Simulation:**
  - https://www.excelschools.net/sims/html/molecule-shapes/latest/molecule-shapes_all.html?locale=en
  - Click 'Model'.
  - Click 'Remove All' to begin.
  - Add one single bond and slowly add more single bonds - only click 'Show Bond Angles' after the shape has been shown! *(Fill out the table below as you go!)*
  - **Do not play with lone pairs yet!**

Number of Bonding Sites	Molecular Geometry	Bond Angle
1	linear	180°
2	linear	180°
3	trigonal planar	120°
4	tetrahedral	109.5°

- Now click 'Remove All', click 'Show Bond Angles' **add one double bond** and **slowly add more single bonds**.
  - What do you notice about the shape of the molecule compared to when they are all single bonds?

[Same] / [Different] shape

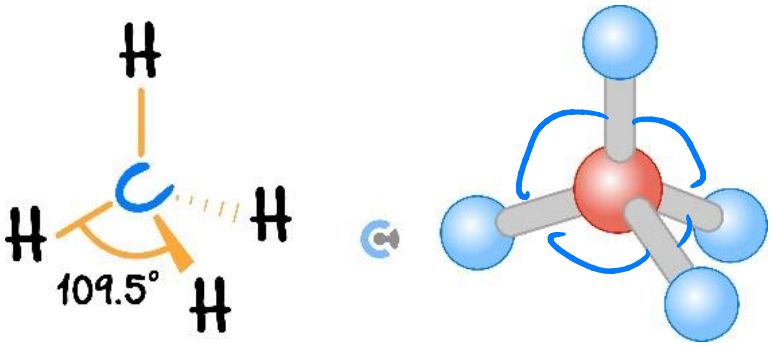
Space for Personal Notes

pair repulsion

- valence shell electron theory VSEPR is used to predict the **shapes** that covalent molecules will take based on **valence electron** _____for molecules.

Number of Bonding Sites	Molecular Geometry	Bond Angle
1	Linear	180°
2	Linear	180°
3	Trigonal planar	120°
4	Tetrahedral	109.5°

NOTE: For molecules with four bonds, consider the following tetrahedral shape:



- The bonds are 109.5° away from each other, which is further away than if they were 90° away!

ALSO NOTE: Only molecules with a maximum of 4 bonds will be tested in VCE Chemistry!

Space for Personal Notes

Try Some Questions!

Question 19 (3 marks)

State the bond angle that exists between bonds in a tetrahedral molecule. Justify why this is the case with reference to a relevant theory.

- 109.5°
- VSEPR Theory
- e^- pairs repel as far as possible in 3D → 109.5°

Question 20 Additional Question.

Which of the following is true?

A. Methane has a trigonal planar molecular geometry.

B. NaCl has a linear geometry.

C. Trigonal planar molecules have an angle of 1200 between their bonds. (120°)

D. The bond angles are random and cannot be predicted for covalent molecules.

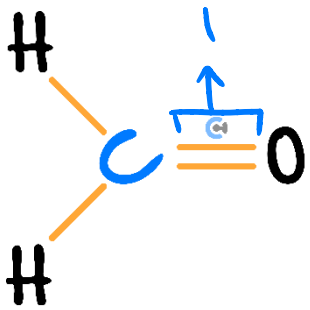
Space for Personal Notes

Sub-Section: Molecular Geometry with Double and Triple Covalent Bonds

How do double and triple bonds affect the shape of a molecule?

Exploration: Methanal

- In double and triple covalent bonds, all electrons associated with the double/triple covalent bonds are considered to be one **electron group.**
- Consider methanal (CH_2O):
- How is the double covalent bond between C and O classified? *(Label Below)*



- In total, how many electron groups are there? 3
- As such, what is the molecular geometry of the molecule? trigonal planar.

Electron Groups

- An electron group is merely a pair of electrons, which may or may not be bonding.
- In double and triple covalent bonds, all electrons associated in those bonds are considered **one** electron group.

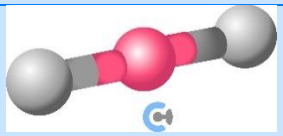
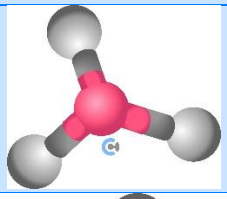
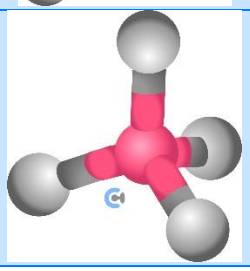
Bonding Site

- A bonding site can be thought of as an **electron group**, provided that group of electrons is part of a bond!

Active Recall: What does VSEPR theory stand for?

valence shell electron pair repulsion

Active Recall: What are the names of each of these shapes, and what are the bond angles?

Number of Single Bonds	Molecular Geometry	Bond Angle	Number of Bonding Sites
	linear	180°	1 - 2
	trigonal planar	120°	3
	tetrahedral	109.5°	4

Let's have a look at a question together!

Question 21 Walkthrough.

C → 4

O → 2

Draw the Lewis Structure of carbon dioxide, and state its molecular geometry.

O = C = O

Molecular Geometry: linear

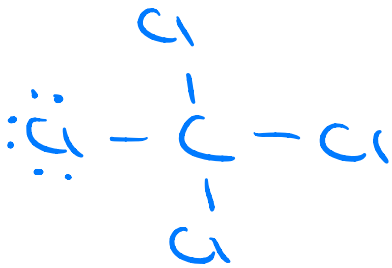
TIP: Regardless as to whether they are single, double or triple bonds, simply just count the number of **electron groups** to determine the molecular geometry (for now)!

Your Turn!

Question 22

Draw the Lewis Structure of each of the following, stating its molecular geometry.

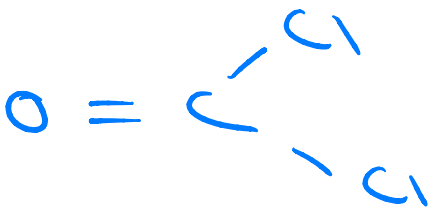
a. Carbon tetrachloride (CCl_4).



C → 4
Cl → 1

Molecular Geometry: tetrahedral

b. Phosgene ($COCl_2$).



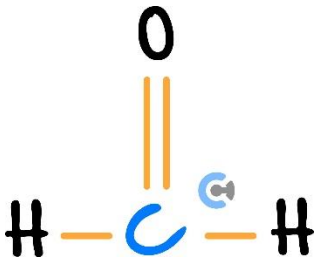
C → 4
O → 2
Cl → 1

Molecular Geometry: trigonal planar.

NOTE: Solid Wedges mean out of the page, Dashed Wedges mean into the page.

Question 23 (3 marks)

Derek is investigating methanal (CHO), and wonders why the bonds are not at 90° to each other like the following:



State the bond angles in the molecule if it were to exist in real life, giving justification for your reasoning.

- 120° → trigonal planar
- 3 bonding site → VSEPR theory causes them to repel → 2D

Question 24 Additional Question.

State the molecular geometry of the following compounds:

a. N_2

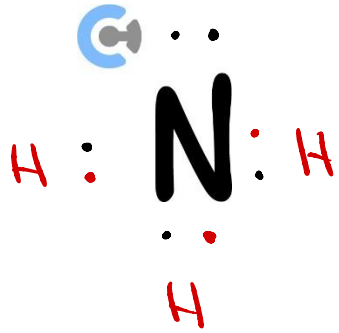
b. HCl

Sub-Section: Parent/Electron Geometry vs ~~Molecular Geometry~~

It's been assumed that all electron groups are bonding sites, but what if there are some non-bonding electrons in the valence shell?

Exploration: Ammonia (NH_3)

- What is the Lewis structure of ammonia? *(Label Below)*
- What do the **bonding** electrons look like in ammonia? *(Label Below)*
- What do the **non-bonding** electrons look like in ammonia? *(Label Below)*



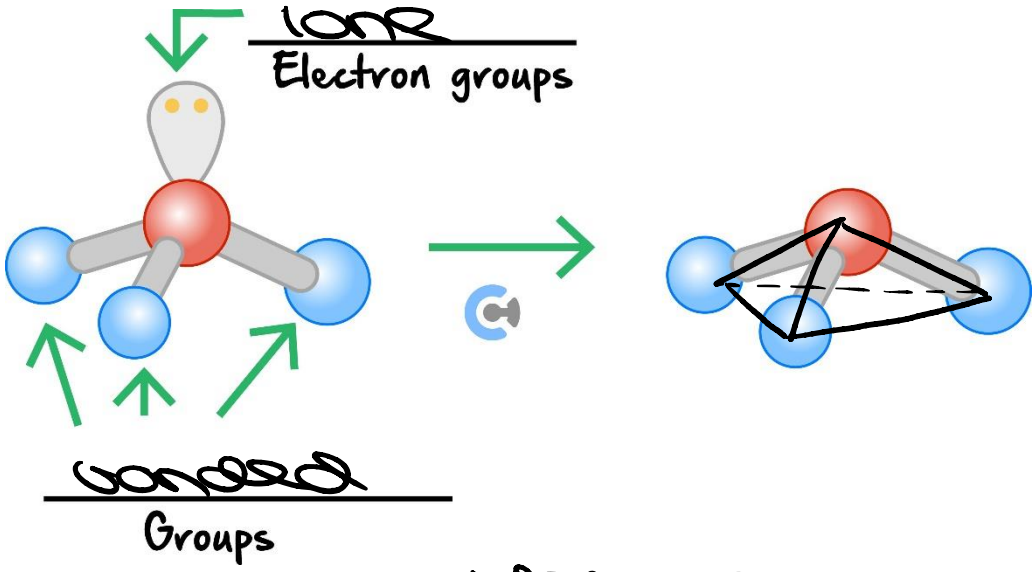
- How many electron pairs/groups are there **in total**? 4
- How many electron pairs/groups are **involved in a bond**? 3
- How many electron pairs/groups are **not involved in a bond**? 1

Lone Pair

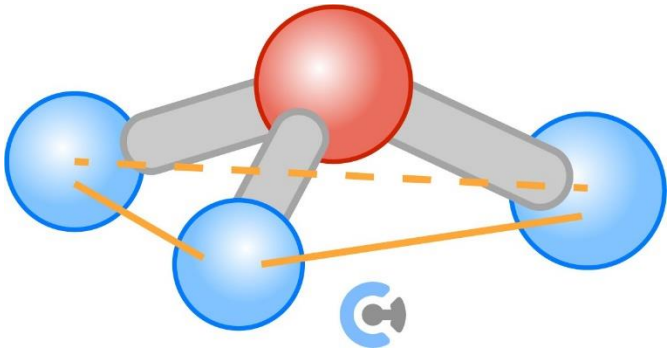
- The non-bonding electron pair is known as the lone pair as it is the pair of electrons which are alone and do not interact with other electrons from other atoms.

Exploration: Parent/Electron Geometry

- **Simulation:**
- https://www.excelschools.net/sims/html/molecule-shapes/latest/molecule-shapes_all.html?locale=en
- Create a molecule with four single bonds. What is the shape?
  - tetrahedral
- Remove one of the single bonds and replace with a lone pair. What shapes do the electrons make?
  - The electron geometry of ammonia (NH_3) is tetrahedral.
- The **shape of the molecule** [is] / [is not] tetrahedral.
- What are the electron groups? *(Label Below)*



- Ammonia (NH_3) takes a pyramidal shape.



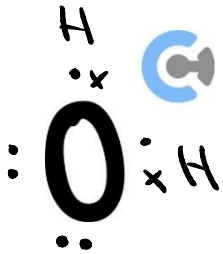
Parent/Electron vs Molecular Geometry

Parent/Electron Geometry	Molecular Geometry
The shape that the arrangement of electron pair in a molecule takes.	The shape that the arrangement of atoms in a molecule takes.
Look at: **total** number of electron pairs (both bonding and non-bonding/lone).	Look at: just number of bonding site bonding electron groups).
Ammonia (NH_3): tetrahedral	Ammonia (NH_3): pyramidal

Let's explore this further!

Exploration: Water (H_2O) Geometry

- If we also consider water (H_2O):
  - What's the Lewis structure of water? *(Label Below)*
  - How are the electron pairs classified? *(Label Below)*



Total Electron Pairs	Bonding Electron Pairs	Non-bonding Electron Pairs	Parent/Electron Geometry	Molecular Geometry
4	2	2	tetra	V-shape

- Construct this on the simulation:
- https://www.excelschools.net/sims/html/molecule-shapes/latest/molecule-shapes_all.html?locale=en

Water Geometry

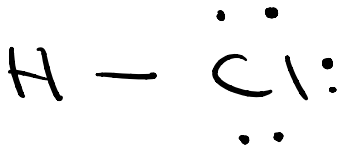
- Water (H_2O) has 2 lone pairs which are not bonded to an atom. It has the following shape:

Parent/Electron Geometry	Molecular Geometry
Tetrahedral	V-shaped/bent

Are there any other shapes?

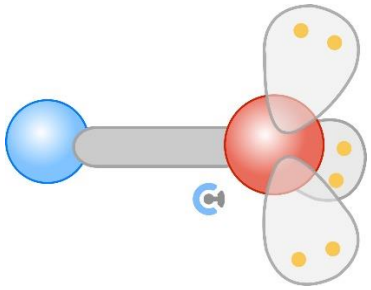
Exploration: Hydrochloric Acid (HCl) Geometry

- Consider the Lewis Structure of HCl:



- Construct this on the simulation: https://www.excelschools.net/sims/html/molecule-shapes/latest/molecule-shapes_all.html?locale=en

- It should look like the following:



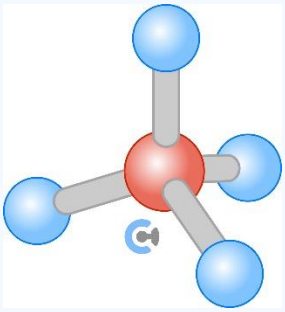
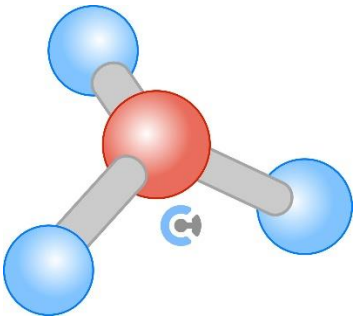
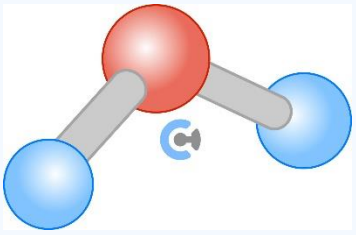
Total Electron Pairs	Bonding Electron Pairs	Non-Bonding/Lone Electron Pairs
4	1	3

- How does the bond look like without the lone pairs? (unclick 'show lone pairs')

Parent/Electron Geometry	Molecular Geometry
tetrahedral	linear

Space for Personal Notes

Parent/Electron Geometry & Molecular Geometry

Number of Bonding electron pairs	Number of Lone Pairs	Molecular Geometry	Shape
4	0	tetrahedral	
3	1	pyramidal	
2	2	V-shape	
1	3	linear	

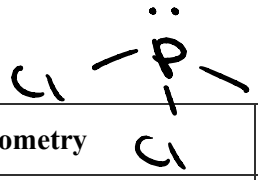
TIP: When figuring out the molecular geometry, first draw the Lewis Structure to determine the total **number of bonding electrons** and then the **number of lone pairs**.

Let's have a look at a question together!

Question 25 Walkthrough.

P → 3
Cl → 1

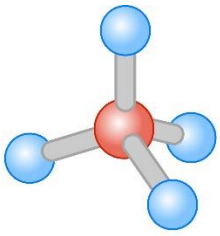
What is the parent geometry and molecular geometry of phosphorous trichloride (PCl_3)?

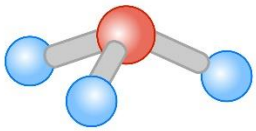


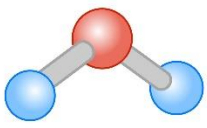
Parent/Electron Geometry	Molecular Geometry
tetrahedral	pyramidal

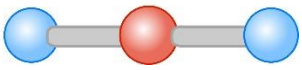
Your Turn!

<u>Active Recall:</u> Name each of the following shapes:

A.  tetrahedral

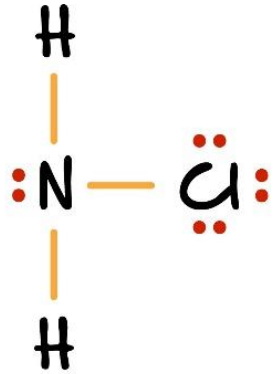
B.  pyramidal

C.  U-shape

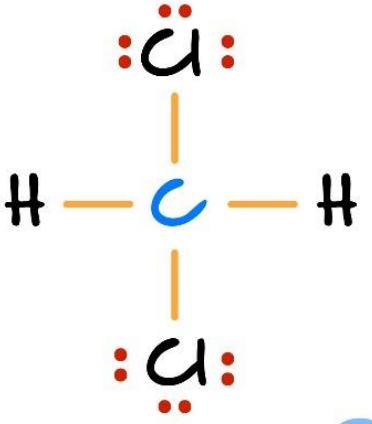
D.  linear

Question 26

Classify the molecular shapes of each of the following atoms:

A. 

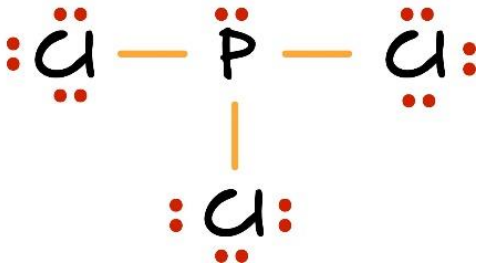
Pyramidal

B. 

tetrahedral

C. 

V-Shaped

D. 

Pyramidal

Question 27

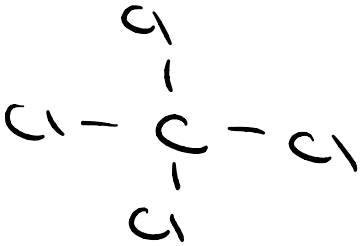
What is the parent geometry and molecular geometry of hydrogen sulphide (H_2S)?

Parent/Electron Geometry	Molecular Geometry
tetrahedral	bent

Question 28

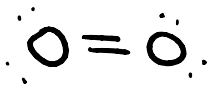
What is the molecular geometry of each of the following?

a. CCl_4



Parent/Electron Geometry	Molecular Geometry
tetrahedral	tetrahedral

b. O_2



Parent/Electron Geometry	Molecular Geometry
trigonal planar	linear

NOTE: The parent/electron geometry is not always tetrahedral!

Question 29 Additional Question.

What is the molecular geometry of each of the following?

a. CH_2ClF

b. HF

c. HOCl

Question 30 Additional Question.

What is the parent geometry and molecular geometry of nitroxyl (HNO)?

Parent/Electron Geometry	Molecular Geometry

Contour Check

Learning Objective: [1.6.1] – Draw Lewis structures of atoms & covalent molecules

Study Design

The use of Lewis (electron dot) structures, structural formulas and molecular formulas to model the following molecules: hydrogen, oxygen, chlorine, nitrogen, hydrogen chloride, carbon dioxide, water, ammonia, methane, ethane and ethene.

Key Takeaways

- The Lewis Structure shows the valence **electrons** in an atom, typically depicted by dots / crosses
- A covalent bond is a chemical bond where two or more atoms [share] / [transfer] electrons.
- [Bonding] / [Non-bonding] electrons are valence electrons directly involved in a covalent bond.
- [Bonding] / [Non-bonding] electrons are electrons **not** directly involved in a covalent bond.
- Covalent bonding occurs between [metals] / [non-metals], as they [gain] / [lose] electrons!
- The molecular formula lists the amount of each atom in a **molecule**.
- Many elements exist in their diatomic form, as the atoms have **full** outer shells
- List the diatomic elements using the mnemonic. *(Label Below)*

  Have no fear of ice cold beer

- The element that can form the [most] / [least] covalent bonds is usually in the **middle** of the molecule.

Learning Objective: [1.6.2] - Identify the geometries (parent & molecular) of molecules, with reference to VSEPR theory

Study Design

Shapes of molecules (linear, bent, pyramidal, and tetrahedral, excluding bond angles) as determined by the repulsion of electron pairs according to valence shell electron pair repulsion (VSEPR) theory.

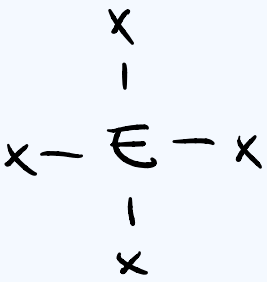
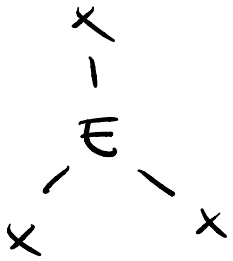
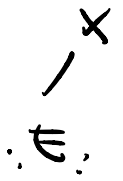
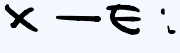
Key Takeaways

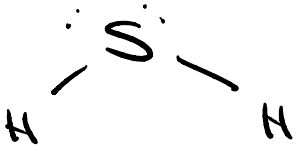
- As electron pairs [attract] / [repel] each other, they will be located as [close] / [far] as possible from each other.
- The _____ **theory** (VSEPR) is used to predict the **shapes** that covalent molecules will take based on valence **electron repulsion** for molecules!

Number of Bonding Sites	Molecular Geometry	Bond Angle
1	linear	180°
2	linear	180°
3	trigonal plane	120°
4	tetrahedral	109.5°

- An electron group is merely a **pair** of electrons, which may or may not be bonding.
- In double and triple covalent bonds, all electrons associated in those bonds are one electron group.
- A bonding group can be thought of as an **electron group**, if the group of electrons is **part of a bond!**
- The non-bonding electron pair is also known as the lone pair as it is the pair of electrons.

Parent/Electron Geometry	Molecular Geometry
The shape that the arrangement of [electrons] / [atoms] in a molecule takes.	The shape that the arrangement of [electrons] / [atoms] in a molecule takes.
Considers **both** the **bonding electron pairs** and the **non-bonding electron pairs** (lone pairs)	Only considers the **atoms**, excluding lone pairs.

Total Number of electron groups	0 non-bonding pairs	1 non-bonding pair	2 non-bonding pairs	3 non-bonding pairs
4				
3				N/A
2			N/A	N/A



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