Learning Guide Module

Subject Code	Chem 1	General Inorganic Chemistry
Module Code	4.0	Chemical Bonding
Lesson Code	4.2	Chemical Structures
Time Limit		30 min

Components	Tasks	AT[a]	ATA[b]
Target	At the end of the chapter, the students must be able to: 1. represent chemical ionic and covalent bonding with Lewis structures 2. illustrate the Lewis structure of different chemical species 3. apply the Octet Rule to determine the chemical bonds between atoms; and 4. explain why there are exceptions to the Octet Rule	1 min	
HOOK	The year 2020 has been a sight marked by turbulent whirlpools of unfortunate events. In the Philippines alone, we've been taken aback by typhoons, volcanic eruptions, earthquakes, and this COVID-19 pandemic, particularly, being the highlight of it all. The pandemic is really a sudden and unprecedented episode for everyone. Going through the first few months of the lockdowns, we faced it head on! With our green thumbs covered in dirt, we did gardening. We became *plantitos* and *plantitas*, carefully watering our succulents. We kneaded dough for cupcakes and pies. We engaged ourselves to online selling. We overemphasized that we can get through this phase with resilience. We thought that we got this. Several months have passed, relationships were tested. Health experts are pretty much concerned that long-term physical distancing could actually trigger mental health issues. Whenever you feel excluded, anxious, or stressed, please know that your perceived emotional states are perfectly valid. Just make sure not to let them consume you totally. As luck would have it, there are eight assured, effective mental health strategies and preemptive actions that one can take to carry on through physical distancing.	3 min	

Fig. 1 *Physical Distancing* [graphic illustration]. (2020).
Retrieved from Imperial College London.

1. **Continue and develop social connections.**
   This includes the old-fashioned phone calls or regular on-screen meetings with family and friends. Reconnect to your relatives that you haven't seen for years. You could also do a casual conversation with your neighbor from a safe distance of two meters.

2. **Be active in the flesh.**
   Exercise can elevate your mood. Go for a walk!

3. **Hush your mind.**
   Have a regular ME-time, meditate, listen to a playlist in *Spotify*, pet a doggo. These reduce anxiety and lower blood pressure.

4. **Go and love yourself.**
   Sleep, eat a balanced diet, and breathe the cool air coming from the open bay. Finish your list on *Netflix*. Have fun and do the things that you love the most.

5. **Disconnect to connect.**
   Limit your screen time on social media for they are the purveyors of bad news that bring mental discomforts. Lean on tales and people that convey positivity and assurance.

6. **Know how to cope up.**
   Draw on skills that can help you get your mind off things. Try cooking a new recipe, learn a new language, read Silberberg's Chemistry book from cover to cover, or pray for better days. This is for you to decide.

7. **Choose kindness.**
   Help others by knowing what they feel. Talk to them to maintain the line of mutually friendly feelings. Count and share your blessings with others.

8. **Seek support and counseling.**
   Ask for help when you need it. Counseling services are available online and can be arranged through various forms of telecommunication. Your school's guidance counselor would be happy to be of service.

Follow these eight tips to help you attain mental stability in the midst of this COVID-19 pandemic. Let us feed our minds with hopeful expectations for these become our inclination of thinking. Stand on guard with your mental health. Keep going! 😊

Ignite

17 min

On the Hook part, you were introduced to eight (8) mental health tips that will help your mind attain stability. Same goes to most of our atoms. They need eight (8) valence electrons to be stable. Once the stability of your mind is achieved, you are ready to go for the stability of the atoms. Let's do this!

LEWIS ELECTRON DOT STRUCTURE (LEDS)

Dots are often used to represent the valence electrons in atoms and molecules. This structure is known as *Lewis structures*, *electron-dot structures*, or *Lewis electron-dot structures*, termed after Gilbert Newton Lewis who is a forerunner in comprehending the connection between electron structure and chemical bonding.

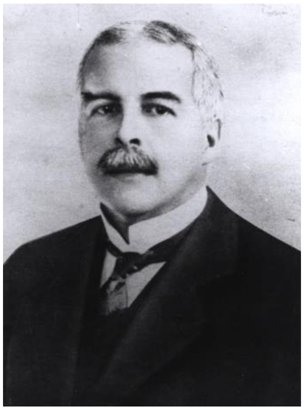


Fig. 2 *Gilbert Newton Lewis* [photograph]. (n.d.).
Retrieved from The Chemist's Club.

This structure consists of the chemical symbol of a particular element surrounded by a number of dots. The nucleus and the inner electrons are represented by the chemical symbol. The dots are in place for the *valence electrons*, electrons found on the outermost shell of the atom.

Table 1. Lewis Electron-Dot Structures of

Some Representative Elements

Group Number							
1A(1)	2A(2)	3A(13)	4A(14)	5A(15)	6A(16)	7A(17)	8A(18)
H •							He:
Li•	Be•	•Ba•	•C•	•N•	•O•	:F•	:Ne:
Na•	Mg•	•Al•	•Si•	•P•	•S•	:Cl•	:Ar:

When writing an electron-dot structure, imagine a square around the symbol of the elements and put a dot on each side until all the valence electrons are used. Once there is one dot on each side, electrons may be doubled up as necessary. The side used for pairing dots is not important. Here's an example for oxygen. Note that oxygen has six valence electrons and its structure could be written in five forms:

1 2 3 4 5

THE OCTET RULE

Elemental sodium is a very active element because it can easily produce compounds with almost all of the nonmetals, with the noble gases as exceptions. Conversely, the noble gas argon resists chemical bonding. Why

The formation of chemical bonds requires changes in the electron configuration of an element. We are all aware that the atoms of noble gases are stable as solidary atoms, a reason leading us to why they do not generally form bonds with other atoms. They have no tendency of changing their electron configurations.

Noble gases, with the exception of He, have completely filled outer *s* and *p* subshells (ns^2np^6). If you noticed, these add up to eight electrons known as an *octet of electrons*.

Octet rule states that the atoms of elements form bonds so as to have access to eight outer electrons. Obtaining this stable configuration is the driving force for bond formation.

The noble gas, helium, forms a stable configuration with only two electrons, a *duet*. This is because helium has a room for two electrons in its only one valence shell. It is a single s orbital, so it can only hold two electrons. Therefore, the element is most stable when its valence orbital is occupied by two electrons. Helium is grouped with elements having eight valence electrons even though it only has two electrons. Well, helium's outermost shell is completely filled making it exceptionally stable. Helium is happy with that.

Fig. 3 *Helium, happy with its two electrons* [graphic illustration]. (n.d.). Retrieved from Shmoop, humor-injected study guides and study tools that make learning easier.

Normally, two electrons pairs up to form a bond just like H_2. Most of the atoms require a maximum of eight electrons in the valence shell to follow the octet structure, CH_4.

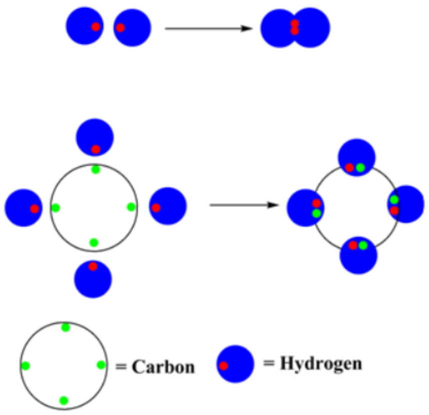


Fig. 4 *Bonding of H_2 and CH_4* [graphic illustration]. (2020). Retrieved from Chem LibreText.

Atoms will react to get in the most stable state as possible, a noble gas configuration. When all orbitals are full, it is stable and the octet is completed.

There are three ways in which an atom can adjust its configuration to obtain a noble gas configuration:

- A metal may give off one to three electrons in order to form a cation. This positive ion has the electron configuration of the preceding noble gas.
- A nonmetal may acquire one to three electrons to form an anion. This negative ion has the electron configuration of the following noble gas.
- Two nonmetals tend to share electrons to achieve the needed number of electrons to satisfy the octet rule.

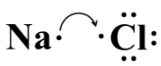
The first two mentioned above talk over the formation of ionic compounds, while the third one brings covalent compounds into being.

How to draw Lewis Diagrams (Ionic Compound)?

Consider the formation of an ionic compound of sodium atom in the presence of a chlorine atom. Their Lewis electron dot diagrams and electron configurations are given below:

$[Ne]3s^1$ $[Ne]3s^23p^5$

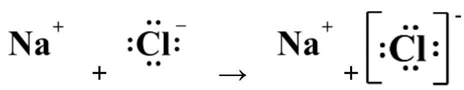
An electron from the Na atom will be transferred to the Cl atom to satisfy the octet.



The resulting ions will be Na^+ and Cl^-. Recall that a neutral atom becomes an ion either by losing an electron (cation) or by gaining an electron (anion).

$[Ne]$ $[Ne]3s^23p^6$

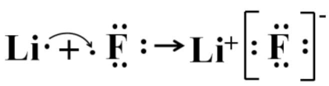
Now, both species have complete octets, and the electron shells are stable. Law of magnetism, we know that opposite charges attract. This is seen in the bond formation of Na^+ and Cl^- ions:



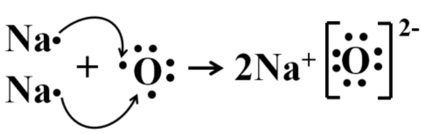
We have written the final formula for sodium chloride as NaCl by doing the simple crisscrossing and dropping of charges. This attraction between atoms bearing opposite charges is called an ionic bond.

Examples:

1. LiF



2. Na_2O



How to draw Lewis Diagrams (Covalent Compound)?

Lewis structures are a suitable way that does not involved much effort of displaying the covalent bonds in many molecules and ions of the representative elements. In writing Lewis structures, connect the atoms in a molecule with covalent bonds by rearranging the valence electrons of the atoms so that each atom has eight outer-shell electrons around it.

Even if the octet rule and Lewis structures do not portray a thorough picture of covalent bonding, they do aid in understanding the ways of bonding in many chemical species and account for the properties of compounds and the chemical reactions they undergo. The following steps will be helpful in learning to write these structures:

1. The total number of valence electrons to be used in the structure must be obtained by simply adding the number of valence electrons of each atom in a molecule or ion. The charge an ion must be taken into account. One electron must be added for each negative charge or one electron must be subtracted for each positive charge on the ion.
2. The skeletal arrangement for the atoms must be written down. They will be connected with a single covalent bond (two dots for one dash).
3. Two electrons will be subtracted for each single bond used in step 2 from the total number of calculated electrons in step 1. This will give you the net number of electrons to spend in order to complete the structure.
4. The pairs of electrons, seen as pairs of dots, must be distributed around each atom except hydrogen. This will give each atom a total of eight electrons, satisfying the octet rule.
5. In some cases, you are required to change the single bonds to double or triple bonds when there are not enough electrons for atoms to attain eight electrons each. This can be done by shifting nonbonding pairs of electron to form multiple bonds as previously stated. Each of the atom must be checked to see if it has eight electrons around it. For hydrogen, it requires only two electrons to satisfy the duet. A double bond is made up of four electrons while a triple bonds counts as six.

Examples:

1. Write the Lewis structure for nitrogen trifluoride (**NF_3**) in which all three F atoms are bonded to the N atom (R. Chang, 2010).

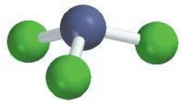


Fig. 5 *NF_3 is a colorless, odorless, unreactive gas.* [graphic illustration]. (2010). Retrieved from Chemistry, 10th Edition.

Step 1: Following the periodic trend in electronegativity, N is less electronegative than F. This makes N the central atom of the compound and F as the surrounding atom. The skeletal structure of NF_3 is:

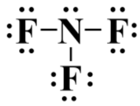
F N F
F

Step 2: Determine the total number of valence electrons of each element present in the formula. In this case, N has 5 and F has 7. The total number of valence electrons is **26**:

N: 5 x 1 = 5

F: 7 x 3 = 21 — Fluorine is multiplied by three because three fluorine atoms are indicated in the formula.

Step 3: Place a line to indicate a single covalent bond between N and each F. The octet for each of the F atom must be completed. The remaining two electrons will be placed on N:



Check: The valence electrons in NF_3 must be counted including the bonds formed and the lone pairs. The result must be equal to 26, the same as the total number of valence electrons on three F atoms (7 x 3 = 21) and one N atom (5 x 1 = 5).

2. Write the Lewis structure for nitric acid (HNO_3) in which the three O atoms are bonded to the central N atom and the ionizable H atom is bonded to one of the O atoms. (R. Chang, 2010).

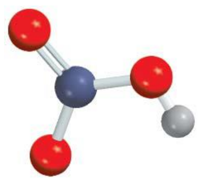


Fig. 6 *HNO_3 is a strong electrolyte* [graphic illustration]. (2010). Retrieved from Chemistry, 10th Edition.

Step 1: Write the skeletal structure for HNO_3.

O N O H
O

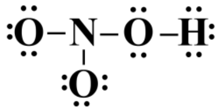
Step 2: Determine the total number of valence electrons of each element present in the formula. In this case, N has 5, O has 6 and H has 1. The total number of valence electrons is **24**:

N: 5 x 1 = 5

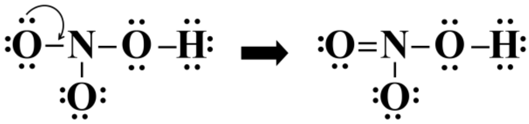
O: 6 x 3 = 18

H: 1 x 1 = 1

Step 3: A single covalent bond will be placed by using a line between N and each of the three O atoms. Another one will be placed between one H atom and the O atom. Distribute the electrons to satisfy the octet rule of each of the O atom:



Step 4: If you can notice, this structure complies with the octet rule for all the O atoms, but not for the N atom. N atom has only six electrons around it. For that reason, a lone pair will be shifted from one O atom for a double bond to be formed with the N atom. Now, check if the octet rule was being followed by each of the atom present:



Check: The valence electrons in HNO_3 must be counted including the bonds formed and the lone pairs. The result must be equal to 24, the same as the total number of valence electrons on one N atom (5 x 1 = 5), three oxygen atoms (6 x 3 = 18), and one hydrogen atom (1 x 1 = 1).

3. Write the Lewis structure for the carbonate ion (CO_3^{2-}).

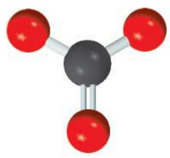


Fig. 7 *Carbonate ion*. [graphic illustration]. (2010).

Retrieved from Chemistry, 10th Edition.

Step 1: We know that C atom is less electronegative than O, making C as the central atom.

O C O
O

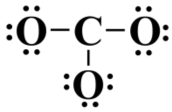
Step 2: Determine the total number of valence electrons of each element present in the formula. In this case, C has 4 and O has 6. Add 2 more electrons because the ion itself has a 2- charge. The total number of valence electrons is **24**:

C: 4 x 1 = 4

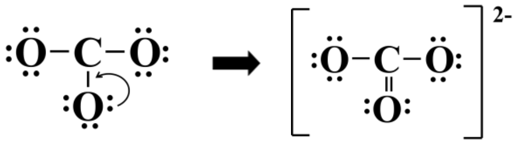
O: 6 x 3 = 18

charge: = 2

Step 3: Place lines in between the atoms and distribute the electrons.



Step 4: If you can notice, this structure complies with the octet rule for all the O atoms, but not for the C atom. C atom here acquired only six electrons. For that reason, a lone pair will be shifted from an O atom to C atom to form a double bond. Now, check if the octet rule was being followed by each of the atom present :



Check: The valence electrons in CO_3^{2-} must be counted including the bonds formed, the lone pairs, and the charge. The result must be equal to 24, the same as the total number of valence electrons on one C atom (4 x 1 = 4), three oxygen atoms (6 x 3 = 18), and the 2- charge.

FORMAL CHARGE AND LEWIS STRUCTURE

If you have observed, a single compound can have different Lewis structures. By evaluating each of the electron involved in a molecule, we can determine and draw the most plausible Lewis structure. This will be discussed in another *Learning Guide Module* about *Resonance*.

A *formal charge* (FC) compares the number of electrons around a "neutral atom" (an atom not in a molecule) versus the number of electrons around an atom in a molecule. Formal charges are being designated to an atom in a molecule by supposing that every electron in all chemical bonds are being shared equally among atoms, without considering their relative electronegativity (D. Kennepohl 2020).

In simpler sense, formal charge is the charge of the different Lewis structures would have if the bonding electrons have equal sharing. To assign the number of electrons on an atom in a Lewis structure, we follow these rules:

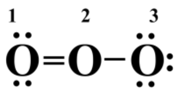
- Every nonbonding electron of an atom is assigned to the atom.
- We break the bond(s) between the atom and other atom(s) and assign half of the bonding electrons to the atom (R. Chang, 2010).

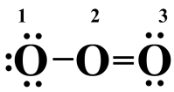
Here's the formula for formal charge:

FC = no. of valence e$^-$ - (no. of unshared e$^-$ + ½ no. of bonding e$^-$)

Example:

Ozone, a gaseous form of oxygen with three oxygen atoms per molecule formed by electrical discharge in oxygen, has a chemical formula of O_3.





A **B**

The formal charge of oxygen 1 in resonance form A is

FC = 6 valence e$^-$ - (4 unshared e$^-$ + ½ of 4 bonding e$^-$)

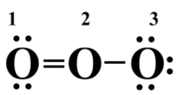
FC = 6 - (4 + 2) = 0

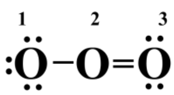
As mental aid, we can modify the formula for formal charges as:

FC = no. of valence e$^-$ - (no. of dots + no. of lines)

The formal charges of all the atoms in the two O_3 Lewis structures are:

	A		**B**
$O_{(1)}$	[6 – (4 + (½·4)] = 0	$O_{(1)}$	[6 – (6 + (½·2)] = -1
$O_{(2)}$	[6 – (2 + (½·6)] = +1	$O_{(2)}$	[6 – (2 + (½·6)] = +1
$O_{(3)}$	[6 – (6 + (½·2)] = -1	$O_{(3)}$	[6 – (4 + (½·4)] = 0





0 +1 -1 -1 +1 0

A **B**

Structures A and B have the same formal charges but on different O atoms, so they contribute equally to the *resonance hybrid*. This will be discussed on a separate Learning Guide Module.

The sum of formal charges must be equal to the actual charge on the species: zero for a neutral molecule or the charge for an ion.

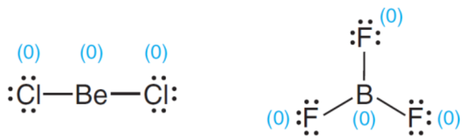
Why is knowing formal charge important? Electrons are involved in formal charges. Keeping them on track will help us establish and predict the reactivity of substances when they undergo chemical reactions.

EXCEPTIONS TO THE OCTET RULE

Exceptions to the octet rule are classified into three categories. These include chemical species having an incomplete octet, an odd number of electrons, and the expanded octet.

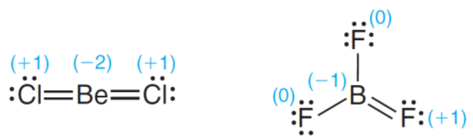
The Incomplete Octet

When either beryllium or boron is assigned as the central atoms of gaseous molecules, they are electron deficient. There are less than eight electrons surrounding the central atom. Study the Lewis structures of gaseous beryllium chloride and boron trifluoride below.

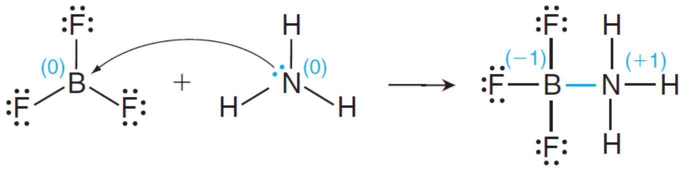


If you can see, there are only four electrons around Be and six around B. Halogens acting as surrounding atoms do not form multiple bonds with the central atom to satisfy the octet rule.

These two structures are most unlikely to happen. You will learn more about this when we encounter *Resonance*.

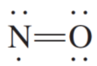


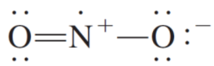
Atoms lacking electrons often attain an octet by having multiple bonds during chemical reactions. For example, boron attains an octet during the reaction of BF_3 with ammonia:



Odd-Electron Molecules

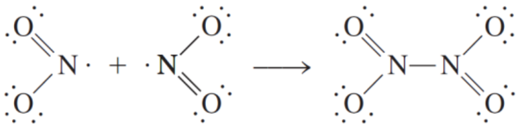
There are molecules that contain odd numbers of electrons. Take nitric oxide (NO) and nitrogen dioxide (NO_2) as examples:





An even number of electrons is needed for pairing completion and that is to reach eight. Clearly, the octet rule won't be satisfied. These molecules are called *free radicals*. These are chemical species that contain an unpaired electron making them extremely reactive and paramagnetic (M. Silberberg, 2015). Free radicals can disrupt genes and membrane. They are also linked to cancer and ageing. That is why taking antioxidants, like vitamin E, is extremely necessary to counteract the action of free radicals.

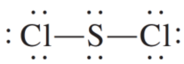
Because of their reactivity, there is a general inclination for the unpaired electron to start forming a covalent bond with an unpaired electron of another molecule. Look at the reaction of two nitrogen dioxides to each other. As you can see, the octet for both N's is satisfied



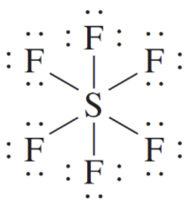
The Expanded Octet

Atoms of elements found in and past the third period of the periodic table can produce compounds that involve more than eight electrons to surround the central atom. These elements can fill their d-orbitals and are used in bonding. This enables them to achieve an expanded octet.

Sulfur obeys the octet rule as seen in sulfur dichloride, SCl_2. Eight electrons surrounds the S atom.



In the case of sulfur hexafluoride, a stable compound, each of its six valence electrons forms a covalent bond with chlorine.



Electronic configuration: $[Ne]3s^2 3p^4$

Twelve electrons surround the central atom, S.

WHY IS OCTET RULE IMPORTANT?

Just like us, humans, chemical species want to attain stability. The octet rule is important because every atom tries hard to achieve full valence shell, the same as the noble gases for they have the most stable electron arrangements.

Navigate

NON-GRADED FORMATIVE ASSESSMENT

I. LEDS: Draw the Lewis structure for the following chemical species:

1. PCl_2^-
2. CH_4O
3. NH_4^+
4. $CaCl_2$
5. MgO

II. LEDS and Formal Charges: Give the three possible Lewis structures of cyanate ion (OCN^-) and give the formal charges of each atom in each structure.

GRADED FORMATIVE ASSESSMENT *(this will be done outside the 30-minute time intended for this module.)*

Read the following items with great comprehension. Give what is asked.

1. The thiocyanate ion (SCN^-), is used in press marks and as a substance that slows down corrosion process when subjected against acidic gases, has at least two possible Lewis electron structures. Give two possible Lewis structures and assign the formal charge of each atom involved. **(6 pts)**
2. Calculate the formal charge of each atom in the NH_4^+ ion. **(5pts)**
3. Draw the Lewis dot structure for the polyatomic ion, PO_4^{3-}. **(2pts)**
4. Draw the Lewis dot structure for formation of the following ionic compounds: **(2pts)**
   a. MgS
   b. AlN

8 min

- The number of valence electrons of an atom is shown in a Lewis dot symbol.
- A Lewis dot structure shows the number of valence electrons possessed by the central atom and its surrounding atoms. It

1 min

Knot 	also shows the bonding pairs of electrons often represented by lines. • A Lewis dot structure shows the placement of atoms in space but do not tell us the actual shape of the chemical species. • The octet rule reflects the observation that main group elements be likely to form bonds in a way that each atom involved will have eight electrons in its outermost shell. This will give them an electron arrangement the same as the noble gases. • Formal charges will help us establish and predict the reactivity of substances when they undergo chemical reactions. This will help us decide which Lewis structure is the most plausible. • Exceptions to the octet rule are classified into three categories. These include chemical species having an incomplete octet, an odd number of electrons, and the expanded octet.		

[a] suggested time allocation set by the teacher

[b] actual time spent by the student (for information purposes only)

References:

Albarico, J. (PSHS-CBRZ). (n.d.). THINK Framework. Based on Science Links by E.G. Ramos and N. Apolinario, Quezon City. Rex Bookstore.

Bayquen, A. (2012). *Chemistry, Exploring Life Through Science*. (G. Bernas PhD, Coordinator.) 2nd Edition. Quezon City, Philippines: Phoenix Publishing House.

Chang R. (2010). *Chemistry, 10th Edition.* New York, McGraw-Hill Company Inc.

Eight Mental Health Tips to Help Persevere Through Physical Distancing (2020). Retrieved from https://chartwell.com/en/blog/2020/05/8-mental-health-tips-to-help-persevere-through-physical-distancing

Kennepohl, D. (2020). *Formal Charges*. Retrieved from Chemistry LibreTexts.

Mendoza, E., Religioso T. (2001). *You and the Natural World Series, Third Year Chemistry Textbook*. (M. Deauna, Coordinator.) 2nd Edition. Quezon City, Philippines: Phoenix Publishing House.

Silberberg M. (2015). *Chemistry: The Molecular Nature of Matter and Change*. 7th Edition. New York, USA: McGraw-Hill Education.

Spinney, R. (2020). *Lewis Structures*. Retrieved from Chemistry LibreTexts.

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