UNIT 3: IONIC AND COVALENT COMPOUNDS

Chapter 5: Ions and Ionic Compounds

5.1: Ions

Chemical Reactivity: - the tendency for an element or a compound to react with other substances.

Octet Rule: - the tendency for electrons to fill the second and third energy levels (8 valence electrons – for main groups – IA to VIIIA columns) to achieve stability.
- in most cases, this means having the same electron arrangement of the nearest noble gas.
- exceptions to the rule include helium (only 2 electrons to fill the first energy level), and the transition metals.

<table>
<thead>
<tr>
<th>Groups or Families</th>
<th>Chemical Properties</th>
<th>Reasons</th>
</tr>
</thead>
<tbody>
<tr>
<td>Alkali Metals (IA)</td>
<td>very reactive metals</td>
<td>Easily shed 1 valence e(^{-}) to achieve an octet</td>
</tr>
<tr>
<td>Alkaline Earth Metals (IIA)</td>
<td>less reactive than alkali metals</td>
<td>Can shed 2 valence e(^{-}) to achieve an octet</td>
</tr>
<tr>
<td>Halogens (VIIA)</td>
<td>very reactive non-metals</td>
<td>Easily gain 1 valence e(^{-}) to achieve an octet</td>
</tr>
<tr>
<td>Noble Gases (VIIIA)</td>
<td>very stable; all are gaseous state at room temperature</td>
<td>Already has an octet configuration as its valence energy level (always ends in (np^6) – except for helium, which ends in (1s^2))</td>
</tr>
</tbody>
</table>

Ions: - when atoms lose or gain electrons, they attain a positive or negative charge.

1. Cations: - positive charged ions (atoms that lose electrons).
- to name cations → element name follows by the word “ion”

Example 1: Draw the energy level diagram for the following cations.
a. Sodium ion = Na\(^{+}\) (11 p\(^{+}\) and 10 e\(^{-}\))
b. Calcium ion = Ca\(^{2+}\) \((20\ p^+\ and\ 18\ e^-)\)

![Diagram of Calcium Ion]

2. **Anions**: negative charged ions (atoms that gain electrons).
   - to name anions → keep the first part of element name follow by suffix \~ide

**Example 2**: Draw the energy level diagram for the following anions.

a. Chloride = Cl\(^-\) \((17\ p^+\ and\ 18\ e^-)\)

![Diagram of Chloride]

b. Oxide = O\(^2-\) \((8\ p^+\ and\ 10\ e^-)\)

![Diagram of Oxide]
**Transition Metals** (1B to 10B): - groups and periods of metals that can have **varying charges**.
- use **Roman Numerals** as part of their ionic names.

**Example:** Fe$^{3+}$ and Fe$^{2+}$ ions

<table>
<thead>
<tr>
<th>26</th>
<th>55.85</th>
</tr>
</thead>
<tbody>
<tr>
<td>3+, 2+</td>
<td></td>
</tr>
</tbody>
</table>

Fe$^{2+}$ = Iron (II) ion  (Old Name – Ferrous ion)

Fe$^{3+}$ = Iron (III) ion  (Old Name – Ferric ion)

(First Charge Listed is the most common type)

**Chemical Properties of Metals and Non-Metals:**

1. **Metals lose electrons to become positive ions – cations.**
2. **Non-Metals gain electrons to become negative ions – anions.**
3. Hydrogen usually loses an electron to become a H$^+$ ion. However, it can sometimes gain an electron to become H$^-$ (Hydride).
4. **The last column of the Table of Elements does not usually form ions.** These elements are called the **Noble Gases** (Helium, Neon, Argon, Krypton, Xenon, and Radon).

The number of electrons an atom loses or gains depends on which column (vertical) the element is at the Table below.

<table>
<thead>
<tr>
<th>I A</th>
<th>I I A</th>
<th>I I I</th>
<th>I V A</th>
<th>V A</th>
<th>V I A</th>
<th>V I I A</th>
<th>V I I I A</th>
</tr>
</thead>
<tbody>
<tr>
<td>+1</td>
<td>+2</td>
<td>8</td>
<td>16.00</td>
<td>0</td>
<td>-3</td>
<td>-2</td>
<td>-1</td>
</tr>
</tbody>
</table>

**Oxygen**

**Common Ion Charge**

These elements do **NOT** usually form ions

**Transition Metals:** - can have more than one type of charge

The reason that noble gases (column VIIIA) do not form ions is because their outermost shells are filled with the maximum number of electrons allowed. That is why we call this group of elements “noble gases”. They do not form ions because they are stable. Hence we use the word “noble” to describe them. All the other elements form ions because they want to achieve stability like the noble gases. If you observe carefully, oxide has the same number of electrons as the nearest noble gas, neon. On the other hand, calcium ion has the same number of electrons as the nearest noble gas, argon. In terms of stability, which is another word for lower energy state, these ions are more stable than their respective atoms.
Since the number of valence electrons of an atom is the same as its column number, all the elements of column IA have 1 valence electron. As we see with lithium, all they have to do is to lose that valence electron to achieve a noble gas “like” state. For elements in column IIA, they all have 2 valence electrons. Hence, they lose 2 electrons to acquire stability and become ions with a net charge of +2. The following table summarises these points.

<table>
<thead>
<tr>
<th>Column</th>
<th>Number of Valence Electrons</th>
<th>Methods to achieve a Stable State</th>
<th>Net Charge of Ions</th>
</tr>
</thead>
<tbody>
<tr>
<td>IA</td>
<td>1</td>
<td>lose 1 electron or gain 7 electrons</td>
<td>+1</td>
</tr>
<tr>
<td>IIA</td>
<td>2</td>
<td>lose 2 electrons or gain 6 electrons</td>
<td>+2</td>
</tr>
<tr>
<td>IIIA</td>
<td>3</td>
<td>lose 3 electrons or gain 5 electrons</td>
<td>+3</td>
</tr>
<tr>
<td>IVA</td>
<td>4</td>
<td>lose 4 electrons or gain 4 electrons</td>
<td>+4</td>
</tr>
<tr>
<td>VA</td>
<td>5</td>
<td>lose 5 electrons or gain 3 electrons</td>
<td>−3</td>
</tr>
<tr>
<td>VIA</td>
<td>6</td>
<td>lose 6 electrons or gain 2 electrons</td>
<td>−2</td>
</tr>
<tr>
<td>VIIA</td>
<td>7</td>
<td>lose 7 electrons or gain 1 electron</td>
<td>−1</td>
</tr>
<tr>
<td>VIIIA</td>
<td>8</td>
<td>already has the maximum number of electrons allowed in the outermost electron shell.</td>
<td>0</td>
</tr>
</tbody>
</table>

**Ionic Equation**: - a chemical equation that shows non-metal elements gaining electrons to become anions, or metal elements losing electrons to become anions,

**Electron Configurations of Ions**: - for metals, which like to lose electrons to form cations, they have the same electron configurations as the noble gas of the previous row. 
- for non-metals, which like to gain electrons to form anions, they have the same electron configurations as the noble gas at the end of the same row.

**Alkali Cations**: - cations that were the result as alkali metals (Group 1 or IA) losing one valence electron. 
- they are Li⁺, Na⁺, K⁺, Rb⁺, Cs⁺ and Fr⁺

**Alkaline Cations**: - cations that were the result as alkaline earths (Group 2 or IIA) losing two valence electrons.
- they are Be²⁺, Mg²⁺, Ca²⁺, Sr²⁺, Ba²⁺ and Ra²⁺

**Halide Ions**: - anions that were the result as halogen (Group 17 or VIIA) gaining one valence electron to satisfy the octet rule.
- they are F⁻, Cl⁻, Br⁻, I⁻ and At⁻

**Example 3**: Write the electron configuration and an ionic equation showing the gaining or losing of electrons when each of the following atoms becomes their most popular ions.

a. sodium atom to sodium ion

```
Na 1s² 2s² 2p⁶ 3s¹ → Na⁺ 1s² 2s² 2p⁶ + e⁻
```
b. aluminium atom to aluminium ion

\[ \text{Al}^{1s^2 2s^2 2p^6 3s^2 3p^1} \rightarrow \text{Al}^{3+} + 3e^- \]

\[ 1s^2 2s^2 2p^6 \]

\[ 1s^2 2s^2 2p^6 3s^2 3p^1 \]

c. fluorine atom to fluoride

\[ \text{F} \rightarrow \text{F}^- \]

\[ 1s^2 2s^2 2p^5 \]

\[ 1s^2 2s^2 2p^6 \]

d. sulfur atom to sulfide

\[ \text{S} + 2e^- \rightarrow \text{S}^{2-} \]

\[ 1s^2 2s^2 2p^6 3s^2 3p^4 \]

\[ 1s^2 2s^2 2p^6 3s^2 3p^6 \]

**Ions Sizes:** - the size of cations and anions.

**Several Notes on Trends in Ion Size**

1. **Metal Cations are generally Smaller than Non-Metal Anions WITHIN the Same Period.** Metal cations tend to lose electrons to achieve the electron configuration of a previous noble gas. Thereby, losing an energy level in the process. **Metal Cations are always Smaller than the Parent Neutral Atoms.** Non-Metal Anions are always Larger than the Parent Neutral Atoms.

2. **In general, Ion Sizes Decrease as one move from LEFT to RIGHT of a period WITHIN the METAL GROUPS and WITHIN the NON-METAL GROUPS.** This is because the increases in protons in the same row increase the effective nuclear charge (the protons in the nucleus have more pull on the outer electrons, decreasing shielding), thus drawing these outer electrons closer to the nucleus, decreasing in sizes as the result.

3. ** Ionic Radii INCREASES Down a Group.** This is due to the fact there are more orbitals as the number of row increases. The outer electrons are, of course, further away from the nucleus.
Example 4: Order the following ions from the smallest to the largest.

a. Te\(^{2-}\), S\(^{2-}\), Se\(^{2-}\), O\(^{2-}\)

These anions are within the same Group (column). As we move down the column, ion size increases. Therefore,

O\(^{2-}\) < S\(^{2-}\) < Se\(^{2-}\) < Te\(^{2-}\)

b. Na\(^+\), Mg\(^{2+}\), P\(^{3-}\), S\(^{2-}\), Cl\(^-\)

These ions are within the same Period (row). As we move to the right, ion size decreases within each of the metal and non-metal groups. Therefore,

Mg\(^{2+}\) < Na\(^+\) < Cl\(^-\) < S\(^{2-}\) < P\(^{3-}\)

Assignment

5.1 pg. 165 #1 to 13; pg. 141 #8 and 13

5.2: Ionic Bonding and Salts

**Ionic Compound:** - when a metal element combines with a non-metal element.
- forms **ionic bonds** (electrons are “stolen” or “transferred” from one atom to another).
- dissociates into **electrolytes** (forms ions, cations and anions, when dissolve in water).
- the resulting ionic compounds are commonly referred to as “salt”.

**Example:** LiF

\[
\begin{array}{c}
\text{Li} + \text{F} \\
\text{valence electrons} \\
\text{Li}^+ \quad \text{F}^- \\
\text{Ionic Bond}
\end{array}
\]

**Lattice Energy:** - the energy involved when one mole of an ionic solid is formed from separated gaseous ions.
- lattice energy is usually exothermic (energy is released).

\[ M^+ (g) + X^- (g) \rightarrow MX (s) \quad \Delta E_{\text{lattice}} < 0 \]

(where M\(^+\) = Metal Cation and X\(^-\) = Non-Metal Anion)

**Determining the Heat to Form Ionic Compounds**

1. **Enthalpy (Heat) Sublimation of Solid Metal into Gaseous Metal (\(\Delta E_{\text{sub}}\)).** \( M (s) \rightarrow M (g) \)
2. **Ionization of Gaseous Metal into Gaseous Cation (I\(_n\)).** \( M (g) \rightarrow M^+ (g) + e^- \)
3. **Dissociation of Non-Metal Molecules into Atoms (\(E_{\text{bond}}\) = Bond Energy).** \( \frac{1}{2} X_2 (g) \rightarrow X (g) \)
4. **Formation of Anion from Non-Metal Atom (\(E_{\text{affinity}}\) = Electron Affinity).** \( X (g) + e^- \rightarrow X^- (g) \)
5. **Formation of Solid Ionic Compound from Gaseous Ions (\(\Delta E_{\text{lattice}}\)).** \( M^+ (g) + X^- (g) \rightarrow MX (s) \)
6. **Summation of all Energies involved for \(\Delta H_{\text{formation}}\).** \( M (s) + \frac{1}{2} X (g) \rightarrow MX (s) \)
Potential Energy Diagram: - diagram that shows the energy changes of various stages of a reaction.

Example: Potential Energy Diagram for the formation of sodium chloride (NaCl).

\[ \text{Na}^+ (g) + \frac{1}{2} \text{Cl}_2 (g) + e^- \rightarrow \text{Na} (s) + \frac{1}{2} \text{Cl}_2 (g) \]

Properties of Ionic Crystalline Solids

1. Ionic Compounds have a definite Crystalline Structure and are Poor Conductors of Electricity and Heat in their Solid Form. Conduction of electricity and heat requires ions to move freely within the solid. The lattice structures of the solid ionic compounds do not allow ions to move freely.

Simple Cubic Unit Crystalline Structure of NaCl

Body Centred Cubic Unit Crystalline Structure of CsCl

Face Centred Cubic Unit Crystalline Structure of ZnS
2. **The smallest unit of an Ionic Compound is called a “Unit Cell”**. Because of the regular crystal structure of an ionic compound, they are not referred to as molecules.

3. **Ionic solids are generally High Melting Points** (typically 300°C to 1000°C). Since a strong force can only shatter the crystal but not bend it as in metals, the energy needed to completely break up the lattice structure (**lattice energy**) is very large and it is the same energy needed to melt the ionic compounds.

4. **Ionic solids are Hard and Brittle**. The lattice structure of all ionic compounds holds the ions in definite positions. When the compound encountered a strong force, the close proximity of the ions stay close together. This causes the crystal to shatter, not bent like metal solid would.

5. **Ionic solids can be Melted to form Liquids that are Electrical Conductors**. Ionic solids melt when the ions gain enough energy to break the lattice structure. They are move freely and can carry electrical charge through the liquid. This explains why a molten ionic substance conducts electricity, but a solid ionic material doesn't. The ions move through the liquid can carry charge from one place to another.

6. **Soluble ionic solids dissolve to form solutions that are Electrical Conductors.** (Not all ionic substances are soluble in water.) Soluble ionic compounds form electrolytes (ions in aqueous from) that allow the conduction of electricity.

### Assignment

5.2 pg. 175 #1 to 6, 8 to 11

### 5.3: Names and Formulas of Ionic Compounds

**Monoatomic Ions**: - ions that came from a single atom (include metal cations and non-metal anions).  
- monoatomic anion ends with suffix ~ide.  
- some transition metal ions require Roman Numeral when there can be more than one type of charge.

**Examples**:  
Na⁺ = sodium ion,  Cl⁻ = chloride,  Pb⁴⁺ = lead (IV) ion,  Zn²⁺ = zinc ion

**Polyatomic Ions**: - ions that contain many atoms. 
- mostly anions (except NH₄⁺ = ammonium ion). 
- most ends with suffixes ~ate or ~ite (some ends with suffix ~ide). 
- some ends with suffix ~ide still (cyanide - CN⁻ and hydroxide - OH⁻).

**Examples**:  
CO₃²⁻ = carbonate,  Cr₂O₇²⁻ = dichromate,  OH⁻ = hydroxide,  SO₃²⁻ = sulfite

**Nomenclature**: - a naming system.

**IUPAC**: - International Union of Pure and Applied Chemistry.

- an organization that oversees the standard regarding chemistry including chemical nomenclature.

**Nomenclature of Ionic Compounds**

1. **Balance the Cation and Anion Charges.**
2. Use **brackets** for multiple Polyatomic Ions.
3. When naming, use ~ide for the non-metal anions.
4. **Metals** that can have **two or more different charges** must use **Roman Numerals** in the names.
Example 1: Write the chemical formula of the followings.

a. sodium chloride

\[ \text{Na}^+ \text{ and } \text{Cl}^- \Rightarrow \text{NaCl} \]

Need 1 \( \text{Na}^+ \) & 1 \( \text{Cl}^- \) to balance charges

b. calcium fluoride

\[ \text{Ca}^{2+} \text{ and } \text{F}^- \Rightarrow \text{CaF}_2 \]

Need 1 \( \text{Ca}^{2+} \) & 2 \( \text{F}^- \) to balance charges

c. ammonium sulfate

\[ \text{NH}_4^+ \text{ and } \text{SO}_4^{2-} \Rightarrow (\text{NH}_4)_2\text{SO}_4 \]

2 \( \text{NH}_4^+ \) & 1 \( \text{SO}_4^{2-} \) to balance charges

d. magnesium hydroxide

\[ \text{Mg}^{2+} \text{ and } \text{OH}^- \Rightarrow \text{Mg(OH)}_2 \]

Need 1 \( \text{Mg}^{2+} \) & 2 \( \text{OH}^- \) to balance charges

e. tin (IV) sulfite

\[ \text{Sn}^{4+} \text{ and } \text{SO}_3^{2-} \Rightarrow \text{Sn(SO}_3)_2 \]

Need 1 \( \text{Sn}^{4+} \) (IV means 4+ charge) & 2 \( \text{SO}_3^{2-} \) to balance charges

f. aluminium oxide

\[ \text{Al}^{3+} \text{ and } \text{O}^{2-} \Rightarrow \text{Al}_2\text{O}_3 \]

Need 2 \( \text{Al}^{3+} \) & 3 \( \text{O}^{2-} \) to balance charges

Oxoanions: - a series of polyatomic ions that contains different number of oxygen atoms.

- \( \text{SO}_4^{2-} \) sulfate
- \( \text{SO}_3^{2-} \) sulfite
- \( \text{NO}_3^- \) nitrate
- \( \text{NO}_2^- \) nitrite
- \( \text{ClO}_4^- \) perchlorate
- \( \text{ClO}_3^- \) chlorate
- \( \text{ClO}_2^- \) chlorite
- \( \text{ClO}^- \) hypochlorite

Example 2: Name the following ionic compounds.

a. \( \text{Na}_2\text{S} \)

\[ \text{Na}^+ \text{ = sodium} \]
\[ \text{S}^2- \text{ = sulfide} \]

sodium sulfide

b. \( \text{KMnO}_4 \)

\[ \text{K}^+ \text{ = potassium} \]
\[ \text{MnO}_4^- \text{ = permanganate} \]

potassium permanganate

c. \( \text{Li}_2\text{S}_2\text{O}_3 \)

\[ \text{Li}^+ \text{ = lithium} \]
\[ \text{S}_2\text{O}_3^{2-} \text{ = thiosulfate} \]

lithium thiosulfate

d. \( \text{CuCr}_2\text{O}_7 \)

\[ \text{Cu}^{2+} \text{ = copper (II)} \]
\[ \text{Cr}_2\text{O}_7^{2-} \text{ = dichromate} \]

copper (II) dichromate

e. \( \text{Fe}_3\text{(BO}_3)_2 \)

\[ \text{Fe}^{2+} \text{ = iron (II)} \]
\[ \text{BO}_3^{3-} \text{ = borate} \]

iron (II) borate

f. \( \text{Ti(ClO}_3)_3 \)

\[ \text{Ti}^{3+} \text{ = titanium (III)} \]
\[ \text{ClO}_3^- \text{ = chlorate} \]

titanium (III) chlorate

Hydrate: - ionic compounds sometimes come with water molecule locked in their crystal form.

- to name hydrates → use the ionic compound name, then write the prefix follow by the word “hydrate”.

Prefixes for Hydrates

<table>
<thead>
<tr>
<th>Prefix</th>
<th>1 – mono</th>
<th>2 – di</th>
<th>3 – tri</th>
<th>4 – tetra</th>
<th>5 – penta</th>
</tr>
</thead>
<tbody>
<tr>
<td>hexa</td>
<td>6 –</td>
<td>7 – hepta</td>
<td>8 – octa</td>
<td>9 – nona</td>
<td>10 – deca</td>
</tr>
</tbody>
</table>

Example: \( \text{CuSO}_4 \cdot 5\text{H}_2\text{O} \)

copper (II) sulfate pentahydrate

Assignment

5.3 pg. 180 (Practice) #1; pg. 180 #1 to 8 and
Worksheet: Nomenclatures of Ionic Compounds
Chapter 5 Review pg. 183–184 #1 to 29
Worksheet: Nomenclatures of Ionic Compounds

Write chemical formula for the compounds below. Show subscript numbers where needed.

1. sodium chloride
2. strontium nitrate
3. calcium fluoride
4. lithium acetate
5. hydrogen phosphate
6. potassium sulfate
7. magnesium chromate
8. calcium hydroxide
9. iron (III) nitrite
10. ammonium sulfate
11. ammonium iodide
12. iron (III) hypobromite
13. lead (II) sulfite
14. copper (I) sulfide
15. aluminum oxide
16. sodium chlorate
17. sodium hydrogen carbonate
18. nickel (II) sulfate heptahydrate
19. silver dichromate
20. barium cyanide
21. titanium (III) periodate
22. aluminum oxalate
23. cobalt (II) borate
24. lithium hydrogen phosphate
25. barium chloride dihydrate
26. iron (II) nitrate hexahydrate
27. nickel (II) thiosulfate
28. vanadium (V) sulfide
29. cadmium phosphide
30. gold (III) thiosulfate

Name the ionic compounds for the chemical formula below. Use Roman numerals where needed.

31. Na₂S
32. Na₂SO₃
33. Na₂SO₄
34. (NH₄)₂CO₃
35. Fe(OH)₃
36. Zn(CH₂COO)₂
37. SnCl₄
38. Ti₂(CrO₄)₃
39. NH₄MnO₄
40. AgNO₂
41. Cu₂O
42. Mn(OOC₃COO)₂
43. Ni₃(BrO₂)₂
44. Sc(IO)₃
45. HgCO₃
46. Cu(NO₃)₂ • 3H₂O
47. Pb(SO₄)₂ • 2H₂O
48. CoF₂ • 4H₂O
49. Zn₃P₂
50. GeO₂

Answers

1. NaCl
2. Sr(NO₃)₂
3. CaF₂
4. LiCH₃COO
5. H₃PO₄
6. K₂SO₄
7. MgCrO₄
8. Ca(OH)₂
9. Al₂(SO₃)₃
10. (NH₄)₂SO₄
11. NH₄I
12. Fe(NO₂)₃
13. PbSO₃
14. Cu₂S
15. Al₂O₃
16. NaClO₃
17. NaHCO₃
18. Ni(BrO₃)₂
19. Ag₂Cr₂O₇
20. Ba(CN)₂
21. Ti(IO₄)₃
22. Al₂(OOC₃COO)₃
23. Co₃(BO₃)₂
24. Li₂HPO₄
25. BaCl₂ • 2H₂O
26. Fe(NO₃)₂ • 6H₂O
27. NiSO₄ • 7H₂O
28. V₂S₅
29. Cd₃P₂
30. Au₂(S₂O₃)₃
31. sodium sulfide
32. sodium sulfate
33. sodium sulfate
34. ammonium carbonate
35. iron (III) hydroxide
36. zinc acetate
37. tin (IV) chloride
38. titanium (III) chromate
39. ammonium permanganate
40. silver nitrite
41. copper (I) oxide
42. manganese (IV) oxalate
43. nickel (II) bromite
44. scandium hypoiodite
45. mercury (II) carbonate
46. copper (II) nitrate trihydrate
47. lead (IV) sulfate dihydrate
48. cobalt (II) fluoride tetrahydrate
49. zinc phosphide
50. germanium oxide
Chapter 6: Covalent Compounds

6.1: Covalent Bonds

**Covalent Bond**: - an attraction force between two non-metal elements where the sharing of valence electrons is taken place so each atom can achieve the stable duet (for hydrogen) or octet (for other representative elements – \(ns^2np^6\)) conditions.

\[
egin{array}{c}
\text{H} \quad + \quad \text{H} \\
\text{Cl} \quad + \quad \text{Cl}
\end{array}
\rightarrow
\begin{array}{c}
\text{H} \quad \text{H} \\
\text{Cl} \quad \text{Cl}
\end{array}
\]

\(2 \text{ valence } e^- \text{ around each H atom (duet rule)}\)
\(8 \text{ valence } e^- \text{ around each Cl atom (octet rule)}\)

**Covalent Compound**: - forms covalent bonds (electrons are “shared” between atoms). The space these electrons share is referred to as **molecular orbitals**.
- sometimes refers to as **molecular compound**.
- usually forms when a non-metal element combines with a non-metal element.
- forms **non-electrolytes** (do not dissociate into ions when dissolve in water).

**Example**: \(\text{CH}_4\)

**Chemical Bond Model**: - chemical bonds are the result when atoms are arranged to achieve a lower energy state.

**Molar Bond Energy**: - the energy to required breaking or released from forming one mole of a particular chemical bond.
- in general, \(\text{Bond Energy} > 0\) (endothermic – requires energy input) when breaking a bond;
- \(\text{Bond Energy} < 0\) (exothermic – releases energy to the surroundings) when forming a bond.
- the **more energy the reaction is released when a bond is formed, the more stable is the molecule**.

**Bond Length**: - the length of a chemical bond as measured by the **lowest energy level**.

![Chemical Bond Model Diagram](image-url)
**Dipole:** the direction of the charge distribution of a polar molecule. - the length of the vector indicates the strength of the bond polarity, whereas the arrow head indicates the direction of the higher electronegative atom (negatively shifted).

**Polar Covalent Bond:** - a bond where electrons are not evenly distributed, resulting in localized charges at either end.

**Non-Polar Covalent Bond:** - a bond where electrons are evenly distributed, resulting in a non-polar molecule.

**Bond Polarity:** - the difference in electronegativity in a molecule. - the larger the bond polarity, the higher the bond energy, and the stronger is the bond. This is why ionic bonds are very strong and ionic compounds are hard.

1. **Non-Polar Covalent Bonds have No Bond Polarity.**
2. **Polar Covalent Bonds have Some Bond Polarity.**
3. **Ionic Bonds have High Bond Polarity.**

**Example 1:** Determine the bond polarity of Cl₂, HCl, and NaCl. Order them from least to greatest and classify them.

<table>
<thead>
<tr>
<th>Bond Polarity</th>
<th>Bond Energy</th>
<th>Bond Strength</th>
</tr>
</thead>
<tbody>
<tr>
<td>0</td>
<td>↑</td>
<td>↑</td>
</tr>
<tr>
<td>1.0</td>
<td>↑</td>
<td>↑</td>
</tr>
<tr>
<td>2.3</td>
<td>↑</td>
<td>↑</td>
</tr>
</tbody>
</table>

1. **Nonpolar**
   - Cl₂: Pairs of electrons in bond shared equally
   - **Bond Polarity = 0**

2. **Polar**
   - Cl₂: Pairs of electrons in bond closer to Cl than H
   - **Bond Polarity = 1.0**

3. **Ionic**
   - Na⁺: Pair of electrons on Cl; not being shared with Na
   - **Bond Polarity = 2.3**
Metallic Bonds: – the result of the attraction between the electrons in the outermost energy level of each metal atom and all the other atoms in the solid metal.

Iron-Iridium Alloy where atoms are organized in layers

6.2: Drawing and Naming Molecules

Nomenclature of Molecular Compounds
1. **Do NOT use charges to balance subscripts.** Use prefixes to name or write the formula’s subscripts.
2. If the first element has one atom in the molecule, do NOT use mono~ as a prefix.
3. The last element uses the suffix ~ide.

Prefixes for Binary Molecular Compounds
1 – mono  3 – tri  5 – penta  7 – hepta  9 – nona  
2 – di  4 – tetra  6 – hexa  8 – octa  10 – deca

Example 1: Name the following molecular compounds.

a. CO  
   1 Carbon and 1 Oxygen  
   **Carbon monoxide**

b. CO₂  
   1 Carbon and 2 Oxygen  
   **Carbon dioxide**

c. N₂O₄  
   2 Nitrogen and 4 Oxygen  
   **Dinitrogen tetraoxide**

Example 2: Provide the chemical formula for the following compounds.

a. sulfur trioxide  
   1 S and 3 O ⇒ **SO₃**

b. diphosphorus pentaoxide  
   2 P and 5 O ⇒ **P₂O₅**

c. silicon dioxide  
   1 Si and 2 O ⇒ **SiO₂**

Common Names for Some Molecular Compounds (Memorize!)

<table>
<thead>
<tr>
<th>Chemical Formula</th>
<th>Common Name</th>
</tr>
</thead>
<tbody>
<tr>
<td>H₂O</td>
<td>Water</td>
</tr>
<tr>
<td>H₂O₂</td>
<td>Hydrogen Peroxide</td>
</tr>
<tr>
<td>O₃</td>
<td>Ozone</td>
</tr>
<tr>
<td>CH₄</td>
<td>Methane</td>
</tr>
<tr>
<td>C₃H₈</td>
<td>Propane</td>
</tr>
<tr>
<td>NH₃</td>
<td>Ammonia</td>
</tr>
<tr>
<td>CH₃OH</td>
<td>Methanol</td>
</tr>
<tr>
<td>C₅H₁₀O₄</td>
<td>Glucose</td>
</tr>
<tr>
<td>C₁₂H₂₂O₁₁</td>
<td>Sucrose</td>
</tr>
</tbody>
</table>

Note: Do NOT use prefixes for the above common molecular compounds!
Lewis Structure: - sometimes refer to as Lewis Dot Diagram.
- shows electrons of valence electron (s and p orbitals) as dots for all atoms of representative elements (Groups IA to VIIIA).
- max of eight electrons around each atomic symbol, and hence Lewis structure cannot be used on transition metals.

Lewis Structure for Atoms in the First Three Rows of the Periodic Table of Elements

<table>
<thead>
<tr>
<th></th>
<th>Cl</th>
<th>F</th>
<th>O</th>
<th>S</th>
<th>N</th>
<th>P</th>
<th>C</th>
<th>Si</th>
<th>B</th>
<th>Al</th>
<th>Be</th>
<th>Mg</th>
<th>Na</th>
</tr>
</thead>
</table>
| H | ![H Cl F O S N P C Si B Al Be Mg Na Lewis Structure](image)
| He | ![He](image)
| Li | ![Li](image)
| Be | ![Be](image)
| B  | ![B](image)
| C  | ![C](image)
| N  | ![N](image)
| O  | ![O](image)
| F  | ![F](image)
| Ne | ![Ne](image)
| Na | ![Na](image)
| Mg | ![Mg](image)
| Al | ![Al](image)
| Si | ![Si](image)
| P  | ![P](image)
| S  | ![S](image)
| Cl | ![Cl](image)
| Ar | ![Ar](image)

Types of Covalent Chemical Bonds:

1. **Single Covalent Bond**: - covalent bond where there are only **two valence electrons** being shared between two atoms (also refer to as bonding electrons).
   
   Unshared Electron Pairs (Lone Pairs): - pairs of electrons not involving in the covalent bond.

   ![Cl Cl Lone Pairs](image)

   **Structural Formula**: - a Lewis dot diagram of a molecule where the bonding electrons are replaced by a short line.

   ![Cl Cl Structural Formula](image)

2. **Double Covalent Bond**: - covalent bond where there are only **four valence electrons** being shared between two atoms. It is denoted by a double line in the structural formula.

   ![O O O O Double Bond](image)

   ![O O O O Double Bond Structural Formula](image)

3. **Triple Covalent Bond**: - covalent bond where there are only **six valence electrons** being shared between two atoms. It is denotes by a triple line in the structural formula.

   ![N N N N Triple Bond](image)

   ![N N N N Triple Bond Structural Formula](image)
Note: In general, bond length decreases with multiple bonds.

<table>
<thead>
<tr>
<th>Bond and Bond Type</th>
<th>Bond Length (pm)</th>
</tr>
</thead>
<tbody>
<tr>
<td>C – C (single bond)</td>
<td>154</td>
</tr>
<tr>
<td>C = C (double bond)</td>
<td>134</td>
</tr>
<tr>
<td>C ≡ C (triple Bond)</td>
<td>120</td>
</tr>
</tbody>
</table>

Note: In general, bond strength increases with multiple bonds.

### Selected Bond Energies (kJ/mol)

<table>
<thead>
<tr>
<th></th>
<th>Single Bonds</th>
<th>Multiple Bonds</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>H – H</td>
<td>436</td>
<td>C = C</td>
</tr>
<tr>
<td>H – F</td>
<td>569</td>
<td>C ≡ C</td>
</tr>
<tr>
<td>H – Cl</td>
<td>431</td>
<td></td>
</tr>
<tr>
<td>H – Br</td>
<td>363</td>
<td></td>
</tr>
<tr>
<td>H – I</td>
<td>295</td>
<td></td>
</tr>
<tr>
<td>C – H</td>
<td>414</td>
<td></td>
</tr>
<tr>
<td>C – C</td>
<td>347</td>
<td></td>
</tr>
<tr>
<td>C – N</td>
<td>389</td>
<td></td>
</tr>
<tr>
<td>C – O</td>
<td>358</td>
<td></td>
</tr>
<tr>
<td>C – F</td>
<td>485</td>
<td></td>
</tr>
<tr>
<td>C – Cl</td>
<td>339</td>
<td></td>
</tr>
<tr>
<td>C – Br</td>
<td>276</td>
<td></td>
</tr>
<tr>
<td>C – I</td>
<td>240</td>
<td></td>
</tr>
<tr>
<td>C – S</td>
<td>259</td>
<td></td>
</tr>
</tbody>
</table>

### Procedure to draw Lewis Dot Diagram for Covalent Molecules

1. Add up all the valence electrons available from all atoms and extra charges from complex ions (ex: charge of 1– means one more electron; charge of 1+ means one less electron).

2. Decide on the central atom. Usually, it is the atom with the most available bonding sites (carbon and silicon has the most bonding sites – 4 valence e⁻) or it is element with the least number of atoms. Place a pair of electrons to form a single bond between each atom.

3. Arrange any remaining electrons around all atoms such that the octet rule is followed for the second and third periods atoms and duet rule is satisfied for the first period atoms. This may sometimes mean setting up multiple bonds (moving lone pairs to the central atom so that it satisfies the octet rule).

4. Final Lewis structure of Polyatomic Ion must include a big square bracket and the indication of net charge.
Example 3: Draw the Lewis structure for the following molecules and polyatomic ions.

a. \( \text{H}_2\text{O} \)

\[
\begin{align*}
2e^- &\quad (2 \text{ H atoms}) \\
+ 6e^- &\quad (1 \text{ O atom}) \\
8e^- &\quad \text{Total} \\
- 4e^- &\quad (2 \text{ single bonds}) \\
4e^- &\quad \text{Left} \\
- 4e^- &\quad (\text{O needs } 4e^- \text{ to achieve octet}) \\
0e^- &\quad 
\end{align*}
\]

Reason for why the \( \text{H}_2\text{O} \) structure is bent will be explained in section 6.3

\[
\text{H} - \text{O} - \text{H}
\]

b. \( \text{CO}_2 \)

\[
\begin{align*}
4e^- &\quad (1 \text{ C atom}) \\
+ 12e^- &\quad (2 \text{ O atoms}) \\
16e^- &\quad \text{Total} \\
- 4e^- &\quad (2 \text{ single bonds}) \\
12e^- &\quad \text{Left} \\
- 12e^- &\quad (\text{each O need } 6e^- \text{ to achieve octet}) \\
0e^- &\quad 
\end{align*}
\]

There are only 4 \( e^- \) around C. Oxygen atoms need to contribute unshared \( e^- \) pairs (2 lone pairs) to form two double bonds.

\[
\text{O} = \text{C} = \text{O}
\]

c. \( \text{NH}_4^+ \)

\[
\begin{align*}
5e^- &\quad (1 \text{ N atom}) \\
+ 4e^- &\quad (4 \text{ H atoms}) \\
- 1e^- &\quad (+1 \text{ charge of NH}_4^+ \text{ ion}) \\
8e^- &\quad \text{Total} \\
- 8e^- &\quad (4 \text{ single bonds}) \\
0e^- &\quad 
\end{align*}
\]

\[
\text{H} - \text{N} - \text{H} \quad \text{H}^+
\]

d. \( \text{CN}^- \)

\[
\begin{align*}
5e^- &\quad (1 \text{ N atom}) \\
+ 4e^- &\quad (1 \text{ C atom}) \\
+ 1e^- &\quad (-1 \text{ charge of CN}^- \text{ ion}) \\
10e^- &\quad \text{Total} \\
- 2e^- &\quad (1 \text{ single bond}) \\
8e^- &\quad \text{Left} \\
- 6e^- &\quad (\text{N need } 6e^- \text{ to achieve octet}) \\
2e^- &\quad (\text{the rest will go to the C atom}) \\
\end{align*}
\]

There are only 4 \( e^- \) around C. Nitrogen atoms need to contribute unshared \( e^- \) pairs (2 lone pairs) to form a triple bond.

\[
\text{C} \equiv \text{N} \equiv \text{C}^-\text{N}^-
\]
Resonance: - occurs when more than one Lewis structure exists for a molecule.

Resonance Structure: - different variations of a resonance molecule can be drawn using a double headed arrow, ↔, between the diagrams.

Example 4: Draw the Lewis structure for nitrate.

\[
\text{Nitrate} = \text{NO}_3^- \\
5e^- \text{ (1 N atom)} \\
+ 18e^- \text{ (3 O atoms)} \\
+ 1e^- \text{ (−1 charge of NO}_3^- \text{ ion)} \\
24e^- \text{ Total} \\
- 6e^- \text{ (3 single bond)} \\
18e^- \text{ Left} \\
-18e^- \text{ (each O atom need 6e− to achieve octet)} \\
0e^- \\
\]

There will be three possible structures as the double bond can “rotate” around the central N atom.

Exceptions to the Octet Rule

1. In the second row, due to the 2s and the 2p orbitals, C, N, O, and F should always follow the octet rule. Overall, the second period elements will never exceed the octet rule.

2. Boron (B) follows the 6e− rule.

3. In the third period, heavier atoms towards the right of the Table often satisfy the octet rule. This being said, there are times that they can use the empty 3d orbitals to exceed the 8 electrons limit. (Examples: P can have 10e− rule; S and Cl can have 10e− or 12e− rule)

4. Elements in higher row can definitely exceed the octet rule. This sometimes applies to larger noble gases as well. (Examples: Br, I, Xe, Rn can have 10e− or 12e−)

Example 5: Draw the Lewis structure for the following molecules.

a. BCl₃

b. PF₅

c. SCl₆

Assignment

6.2 pg. 202 #1 & 2 (Practice); pg. 203 #1 & 2 (Practice); pg. 205 #1 & 2 (Practice); pg. 207 #1 to 13
6.3: Molecular Shape

Valence Shell Electron-Pair Repulsion (VSEPR) Model:
- the best structure for a molecule is one that minimizes electrons lone pairs repulsion.
- most often used to predict molecular structures involving non-metals.

Example: For molecules with a total of $4\ e^-$ pairs, the bond angles decreases from $109.5^\circ$ as more lone pairs added. (Repulsion of Lone Pair(s) with bond electrons pushed the angle down.)

Effective Electron Pairs: sometimes refer to as substituents.
- the number of lone pairs on the central atom of a molecule and the number of connections between the central atom with the outer atom(s). Each set of Multiple Bond (like double and triple bonds) count as one connection or one effective electron pair.

Summary of Geometrical Shape of Single Bond Molecules

<table>
<thead>
<tr>
<th>Around Central Atom</th>
<th>Total # of Eff. $e^-$ Pairs</th>
<th># of Lone Pairs</th>
<th>Molecular Structures</th>
<th>Geometrical Shape</th>
<th>Example and Bond Angles</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>2</td>
<td>0</td>
<td><img src="image" alt="Linear" /></td>
<td>Linear</td>
<td>BeF$_2$  $\begin{array}{c} \text{F} \ \text{Be} \ \text{F} \end{array}$  $180^\circ$</td>
</tr>
<tr>
<td></td>
<td>3</td>
<td>0</td>
<td><img src="image" alt="Trigonal Planar" /></td>
<td>BCl$_3$  $\begin{array}{c} \text{Cl} \ \text{Cl} \ \text{Cl} \end{array}$  $120^\circ$</td>
<td></td>
</tr>
<tr>
<td></td>
<td>4</td>
<td>0</td>
<td><img src="image" alt="Tetrahedral" /></td>
<td>Tetrahedral</td>
<td>CH$_4$  $\begin{array}{c} \text{H} \ \text{H} \ \text{H} \ \text{H} \end{array}$  $109.5^\circ$</td>
</tr>
<tr>
<td></td>
<td>4</td>
<td>1</td>
<td><img src="image" alt="Trigonal pyramid" /></td>
<td>NH$_3$  $\begin{array}{c} \text{H} \ \text{N} \ \text{H} \end{array}$  $107^\circ$</td>
<td></td>
</tr>
<tr>
<td></td>
<td>4</td>
<td>2</td>
<td><img src="image" alt="V-Shape (Bent)" /></td>
<td>H$_2$O  $\begin{array}{c} \text{H} \ \text{O} \ \text{H} \end{array}$  $104.5^\circ$</td>
<td></td>
</tr>
</tbody>
</table>
### Around Central Atom

<table>
<thead>
<tr>
<th>Total # of Eff. $e^-$ Pairs</th>
<th># of Lone Pairs</th>
<th>Molecular Structures</th>
<th>Geometrical Shape</th>
<th>Example and Bond Angles</th>
</tr>
</thead>
<tbody>
<tr>
<td>5</td>
<td>0</td>
<td><img src="#" alt="Trigonal bipyramid" /></td>
<td>Trigonal bipyramid</td>
<td><img src="#" alt="PF$_5$" /></td>
</tr>
<tr>
<td>5</td>
<td>1</td>
<td><img src="#" alt="See-saw" /></td>
<td>See-saw</td>
<td><img src="#" alt="SF$_4$" /></td>
</tr>
<tr>
<td>5</td>
<td>2</td>
<td><img src="#" alt="T-Shape" /></td>
<td>T-Shape</td>
<td><img src="#" alt="ClF$_3$" /></td>
</tr>
<tr>
<td>5</td>
<td>3</td>
<td><img src="#" alt="Linear" /></td>
<td>Linear</td>
<td><img src="#" alt="Br$_3^-$" /></td>
</tr>
<tr>
<td>6</td>
<td>0</td>
<td><img src="#" alt="Octahedral" /></td>
<td>Octahedral</td>
<td><img src="#" alt="SCl$_6$" /></td>
</tr>
</tbody>
</table>
## Around Central Atom

<table>
<thead>
<tr>
<th>Total # of Eff. e⁻ Pairs</th>
<th># of Lone Pairs</th>
<th>Molecular Structures</th>
<th>Geometrical Shape</th>
<th>Example and Bond Angles</th>
</tr>
</thead>
<tbody>
<tr>
<td>6</td>
<td>1</td>
<td><img src="image1.png" alt="Image" /></td>
<td>Square pyramid</td>
<td>ICl₅</td>
</tr>
<tr>
<td></td>
<td></td>
<td><img src="image2.png" alt="Image" /></td>
<td></td>
<td>Cl Cl Cl Cl Cl</td>
</tr>
<tr>
<td></td>
<td></td>
<td><img src="image3.png" alt="Image" /></td>
<td></td>
<td>Cl</td>
</tr>
<tr>
<td>6</td>
<td>2</td>
<td><img src="image4.png" alt="Image" /></td>
<td>Square planar</td>
<td>XeF₄</td>
</tr>
<tr>
<td></td>
<td></td>
<td><img src="image5.png" alt="Image" /></td>
<td></td>
<td>F F F F F</td>
</tr>
</tbody>
</table>

### Using VSEPR Model with Multiple Bonds Molecule.

1. Look at multiple bonds as single bonds (only to determine molecule geometry).
2. Using the number of bonds and line pairs around the central atom, determine the geometry of the molecule from the charts above.

**Example 1:** From Example 4 of Section 6.2, determine the geometrical shape of NO₃⁻. What is the possible bond angle in this complex ion?

Since there are 3 bonds around nitrogen (we look at the double bond as a single bond when deciding on molecular geometry), and there are no lone pairs around the central atom, the NO₃⁻ ion is in the shape of **trigonal planar** with bond angles at 120°.
Polarity of Polyatomic Molecules: - when dipoles of each bond are treated like vectors and are added up:

- a polar molecule will result when dipoles do NOT cancel each other out.

Example 2: Using Figure 6.20 of the textbook, determine the bond polarity and dipole (if any) of the following molecules.

<table>
<thead>
<tr>
<th>Molecule</th>
<th>Bond Polarity</th>
<th>Overall Dipole</th>
</tr>
</thead>
<tbody>
<tr>
<td>a. CO₂</td>
<td>0.89 (Non-polar)</td>
<td></td>
</tr>
<tr>
<td>b. SO₃</td>
<td>0.86 (Non-polar)</td>
<td></td>
</tr>
<tr>
<td>c. OCl₂</td>
<td>0.28 (Polar)</td>
<td></td>
</tr>
<tr>
<td>d. NF₃</td>
<td>0.94 (Polar)</td>
<td></td>
</tr>
<tr>
<td>e. CH₄</td>
<td>0.35 (Non-polar)</td>
<td></td>
</tr>
<tr>
<td>f. CH₃Cl</td>
<td>0.35 &amp; 0.61 (Polar)</td>
<td></td>
</tr>
</tbody>
</table>
Polarity Affects Boiling Point: - the stronger the polarity of a molecule, the higher its boiling point.
- this is because the polar molecules form positive and negative regions so there is more attraction between molecules (intermolecular forces). Hence, more energy (higher temperature) is required to pull these molecules from one another to change from the liquid phase to the gaseous phase.

<table>
<thead>
<tr>
<th>Boiling Point of Polar Molecules</th>
<th>Boiling Point of Non-Polar Molecules</th>
</tr>
</thead>
<tbody>
<tr>
<td>Water (H₂O)</td>
<td>100°C</td>
</tr>
<tr>
<td>Carbon dioxide (CO₂)</td>
<td>−78.5°C</td>
</tr>
<tr>
<td>Chloroform (CHCl₃)</td>
<td>62.3°C</td>
</tr>
<tr>
<td>Carbon tetrafluoride (CF₄)</td>
<td>−130°C</td>
</tr>
<tr>
<td>Ammonia</td>
<td>−33°C</td>
</tr>
<tr>
<td>Fluorine (F₂)</td>
<td>−188°C</td>
</tr>
</tbody>
</table>

Example 3: Order the boiling points from the least to greatest for the following compounds, PH₃, CH₃F and SiH₄.

PH₃ has a trigonal pyramid geometry (VSEPR) and is Polar. Even though the P–H bonds have no polarity (electronegativities of P and H are the same), the lone pair on one end of the P atom causes an uneven distribution of electrons.

CH₃F has a tetrahedral geometry and is very polar. The C–F bond along with the C–H bonds have strong polarity. The overall dipole for the molecule has electrons around the F atom.

SiH₄ has a tetrahedral geometry with equal dipoles of Si–H bonds cancels out all bond polarities. Hence, SiH₄ is Non-Polar.

Since non-polar molecule have no dipole interactions, SiH₄ should have the lowest boiling point. PH₃ is less polar than CH₃F due to the difference in electronegativities between P–H bond and C–F with C–H bonds. Therefore, CH₃F must have the highest boiling point.

Boiling Point: SiH₄ < PH₃ < CH₃F

Assignment
6.3 pg. 211 (Practice) #1; pg. 213 #1 to 11 and
Worksheet: Molecular Geometry and VSEPR Theory
Chapter 6 Review pg. 216–218 #11 to 40, 42, 44 to 46, 48
Worksheet: Molecular Geometry and VSEPR Theory

1. For the following molecules and polyatomic ions,
   - draw the Electron (Lewis) Dot Diagram (2-Dimensional), show all resonance structures if any;
   - draw the 3-Dimensional Lewis Dot Diagram (with wedged and dotted bond notations);
   - state the Number of Effective Electron Pairs and the Number of Lone Pairs around the central atom;
   - determine the Shape of the Structure and state the bond angles;
   - evaluate (give reasons why) the overall Polarity of the molecule or polyatomic ion.

   a. Br$_2$
   b. CO
   c. BF$_3$
   d. NO$_2^-$
   e. CF$_4$
   f. CHBr$_3$
   g. NF$_3$
   h. SiO$_3^{2-}$
   i. CS$_2$
   j. H$_2$Te
   k. PBr$_5$
   l. SeI$_4$
   m. IF$_3$
   n. BrCl$_2^-$
   o. SBr$_6$

   **Answers**

   a. Br$_2$

   No central atom; Linear (diatomic molecules); Bond Angle = 180°; Non-polar (both atoms have equal electronegativity values)

   b. CO

   No central atom; Linear (diatomic molecules); Bond Angle = 180°; Polar (both atoms have different electronegativity values)

   c. BF$_3$

   Around central atom B (3 eff e$^-$ pairs total, no lone pair); Trigonal Planar; Bond Angles = 120°; Non-polar (no lone pair on central atom and all bonds are equivalent)

   d. NO$_2^-$

   Around central atom N (3 eff e$^-$ pairs total, 1 lone pair); Bent; Bond Angles = less than 120° (about 117.5°); Polar (lone pair on central atom)

   e. CF$_4$

   Around central atom C (4 eff e$^-$ pairs total, no lone pair); Tetrahedral; Bond Angles = 109.5°; Non-polar (no lone pair on central atom and all bonds are equivalent)

   f. CHBr$_3$

   Around central atom C (4 eff e$^-$ pairs total, no lone pair); Tetrahedral; Bond Angles = 109.5°; Polar (no lone pair on central atom but not all bonds are equivalent)

   g. NF$_3$

   Around central atom N (4 eff e$^-$ pairs total, 1 lone pair); Trigonal Pyramid; Bond Angles = 107°; Polar (lone pair on central atom)

   h. SiO$_3^{2-}$

   Around central atom Si (3 eff e$^-$ pairs total, no lone pair); Trigonal Planar; Bond Angles = 120°; Non-polar (no lone pair on central atom and all bonds are equivalent)

   i. CS$_2$

   Around central atom C (2 eff e$^-$ pairs total, no lone pair); Linear; Bond Angles = 180°; Non-polar (no lone pair on central atom and all bonds are equivalent)
Chemistry

Unit 3: Ionic and Covalent Compounds

j. H₂Te

Around central atom Te (4 eff e⁻ pairs total, 2 lone pairs); V-Shaped or Bent; Bond Angles = 105°; Polar (lone pairs on central atom)

k. PBr₅

Around central atom P (5 eff e⁻ pairs total, no lone pair); Trigonal Bipyramid; Bond Angles = 90° and 120°; Non-polar (no lone pair on central atom and all bonds are equivalent)

l. SeI₄

Around central atom Se (5 eff e⁻ pairs total, 1 lone pair); See-Saw; Bond Angles = 88.5° and 118.5; Polar (lone pair on central atom)

m. IF₃

Around central atom I (5 eff e⁻ pairs total, 2 lone pairs); T-shaped; Bond Angles = 88.5° and 183°; Polar (lone pairs on central atom)

n. BrCl₂⁻

Around central atom Br (5 eff e⁻ pairs total, 3 lone pairs); Linear; Bond Angles = 180°; Non-polar (even with lone pairs on central atom, the lone pairs cancelled out with each other and all other bonds are equivalent)

o. SBr₆

Around central atom S (6 eff e⁻ pairs total, 0 lone pair); Octahedral; Bond Angles = 90°; Non-polar (no lone pair on central atom and all bonds are equivalent)

p. ClF₅

Around central atom Cl (6 eff e⁻ pairs total, 1 lone pair); Square-Planar; Bond Angles = 88.5°; Polar (lone pair on central atom)

q. XeBr₄

Around central atom Xe (6 eff e⁻ pairs total, 2 lone pairs); Square Planar; Bond Angles = 90°; Non-polar (even with lone pairs on central atom, the lone pairs cancelled out with each other and all other bonds are equivalent)