

STUDY NOTES IN CHEMISTRY 1

Ionic and Covalent Bonding, Lewis Structures, Polarity, VSEPR, and Intermolecular Forces (Week 7 – 11)

POINTS TO MEMORIZE

1. **NaCl is ionic.**
2. **CO₂ is covalent.**
3. Ionic bonding involves **electron transfer**.
4. Covalent bonding involves **electron sharing**.
5. **Metal + nonmetal → usually ionic.**
6. **Nonmetal + nonmetal → usually covalent.**
7. CaCO₃ has ionic attraction between **Ca²⁺ and CO₃²⁻**, with covalent bonds within CO₃²⁻.
8. The **octet rule** involves achieving a stable valence-shell arrangement.
9. CO₂ is written **O=C=O**.
10. NH₃ has **3 bonds + 1 lone pair**.
11. H₂O has **2 bonds + 2 lone pairs**.
12. CO₂ is **linear and nonpolar**.
13. H₂O is **bent and polar**.
14. NH₃ is **trigonal pyramidal and polar**.
15. VSEPR is based on **repulsion between electron domains**.
16. **London dispersion forces occur in all atoms and molecules.**
17. HCl experiences **dipole-dipole attractions**.
18. H₂O can form **hydrogen bonds**.
19. Stronger intermolecular forces generally mean **higher boiling point**.
20. Boiling generally overcomes **intermolecular forces**, not the covalent bonds within the molecules.

IONIC AND COVALENT COMPOUNDS

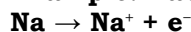
A. Ionic Compounds

What is an ionic compound?

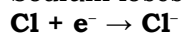
An **ionic compound** forms when electrons are generally **transferred** from one atom to another, producing oppositely charged ions.

- **Metal → loses electrons → positive ion (cation)**
- **Nonmetal → gains electrons → negative ion (anion)**
- The oppositely charged ions attract through **electrostatic forces**.

Example: NaCl



Sodium loses one electron.



Chlorine gains one electron.

Therefore:



Important characteristics

Ionic compounds usually:

- consist of **ions**
- form an **extended crystal lattice**
- have relatively **high melting points**
- do **not** conduct electricity as solids because the ions cannot move freely
- conduct electricity when **molten or dissolved in water** because the ions become mobile

Examples

Compound	Ions	Type
NaCl	Na ⁺ , Cl ⁻	Ionic
MgO	Mg ²⁺ , O ²⁻	Ionic
CaCl ₂	Ca ²⁺ , Cl ⁻	Ionic
CaF ₂	Ca ²⁺ , F ⁻	Ionic

Remember:

Metal + nonmetal → usually ionic

COVALENT COMPOUNDS

What is a covalent compound?

A **covalent compound** forms when atoms **share electrons**.

Covalent bonding commonly occurs between **nonmetal atoms**.

Example: HCl

Hydrogen and chlorine share electrons.



The electrons are shared rather than completely transferred.

Examples

- H₂O
- CO₂
- CH₄
- NH₃
- Cl₂
- O₂

Molecular compounds

Many covalent substances consist of **discrete molecules** rather than an extended ionic lattice.

For example:

CO₂

consists of individual CO₂ molecules.

General comparison

Ionic	Covalent
Electrons transferred	Electrons shared
Usually metal + nonmetal	Usually nonmetal + nonmetal
Forms ions	Forms molecules
Crystal lattice	Discrete molecules are common
Often high melting point	Often lower melting point
Conducts when molten/aqueous	Usually poor electrical conductor

IMPORTANT EXCEPTION: POLYATOMIC IONS

Not every compound can be classified simply by looking at each element.

Example: CaCO₃

Calcium carbonate contains:

Ca²⁺ and CO₃²⁻

There is an **ionic attraction between Ca²⁺ and CO₃²⁻**.

However, the atoms **inside CO₃²⁻ are connected by covalent bonds**.

So CaCO₃ has:

Ionic bonding between the ions + covalent bonding within the polyatomic ion

The same idea applies to compounds such as **NaOH**:

Na⁺ | OH⁻

The attraction between Na⁺ and OH⁻ is ionic, while the O-H bond within OH⁻ is covalent.

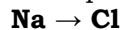
ELECTRON TRANSFER VS. ELECTRON SHARING

This is one of the most important distinctions in the test.

Ionic bonding

Electron transfer

Example:



Na loses an electron and Cl gains it.

Covalent bonding

Electron sharing

Example:



H and Cl share electrons.

Quick clue

Ask:

Are electrons transferred or shared?

- Transferred → ionic
- Shared → covalent

ELECTRONEGATIVITY AND BOND POLARITY

What is electronegativity?

Electronegativity is an atom's ability to attract shared electrons toward itself. Fluorine is the most electronegative element.

When two atoms have similar electronegativities:

The electrons are shared relatively equally.

→ **Nonpolar covalent bond**

Example:



Both chlorine atoms have the same electronegativity.

When one atom is more electronegative:

The shared electrons are pulled closer to that atom.

→ **Polar covalent bond**

Example:



Fluorine attracts the shared electrons more strongly than hydrogen.

Bond polarity examples



Same atoms → same electronegativity

Nonpolar



Same atoms → same electronegativity

Nonpolar



Oxygen is more electronegative than carbon.

Polar



Fluorine is much more electronegative than hydrogen.

Very polar

Important distinction

Bond polarity ≠ molecular polarity

A molecule can contain polar bonds but still be **nonpolar overall**.

LEWIS STRUCTURES (LEDS)

What is a Lewis structure?

A **Lewis structure** shows:

- valence electrons
- bonding electrons
- lone pairs

It helps us understand how atoms bond.

THE OCTET RULE

The **octet rule** states that many atoms tend to achieve **8 electrons in their valence shell**.

Important exceptions

Hydrogen is stable with **2 electrons**, not 8.

HOW TO DRAW LEWIS STRUCTURES

Basic steps

Step 1: Count valence electrons

Example: CH₄

Carbon = 4 valence electrons

Hydrogen = 1 × 4

Total:

4 + 4 = 8 valence electrons

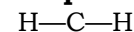
Step 2: Select the central atom

Usually the least electronegative atom is placed in the center.

For CH₄:

C = central atom

Step 3: Connect atoms with single bonds



with two additional H atoms bonded to C.

Step 4: Complete the octets

Carbon forms four bonds.

Each H has two electrons through its bond.

Alam niyo na yan. Nasa ppt ang visuals.

COMMON LEWIS STRUCTURES

CH₄ — Methane

Carbon forms **four single bonds**.

Four C—H bonds

Carbon has a complete octet.

NH₃ — Ammonia

Nitrogen has:

- 3 bonding pairs
- 1 lone pair

Therefore:

3 bonds + 1 lone pair

This is important because the lone pair affects the molecule's shape.

H₂O — Water

Oxygen has:

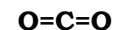
- 2 bonding pairs
- 2 lone pairs

Therefore:

2 O—H bonds + 2 lone pairs

CO₂ — Carbon dioxide

The correct Lewis structure is:



Carbon forms two double bonds.

This gives carbon and oxygen appropriate electron arrangements consistent with the octet rule.

DIATOMIC MOLECULES

Some elements naturally form molecules containing two atoms.

Examples:

- H₂
- O₂
- N₂
- F₂
- Cl₂
- Br₂
- I₂

Example: Cl₂

The two chlorine atoms have the same electronegativity.

Therefore, they **share electrons** rather than transferring an electron completely.

Cl—Cl

The bond is **nonpolar covalent**.

IUPAC NAMING OF IONIC COMPOUNDS

A. Binary ionic compounds

For a simple ionic compound:

metal name + nonmetal ending in -ide

Examples

NaCl

→ **sodium chloride**

MgO

→ **magnesium oxide**

CaCl₂

→ **calcium chloride**

TRANSITION METALS AND ROMAN NUMERALS

Some metals can form more than one charge.

For these metals, the charge is indicated using a **Roman numeral**.

Example: FeCl₃

Each chloride = -1

Three chlorides:

3(-1) = -3

Therefore Fe must be:

Fe³⁺

Name:

iron(III) chloride

Remember

Roman numeral	Charge
I	+1
II	+2
III	+3
IV	+4
V	+5

WRITING IONIC FORMULAS

The total charge of an ionic compound must equal **zero**.

Example: Calcium nitrate

Calcium:

Ca²⁺

Nitrate:

NO₃⁻

We need two nitrate ions to balance one Ca²⁺:

Ca(NO₃)₂

Why parentheses?

Because there are **two complete nitrate ions**.

Example: Aluminum oxide

Aluminum:

Al³⁺

Oxygen:

O²⁻

The smallest combination that gives zero total charge is:

2 Al³⁺ = +6

3 O²⁻ = -6

Therefore:

Al₂O₃

Do you need to memorize all oxidation numbers? Hindi na. Yung common lang.

NAMING COVALENT COMPOUNDS

Covalent compounds commonly use **prefixes** to indicate the number of atoms.

Number	Prefix
1	mono-
2	di-
3	tri-

Number	Prefix
4	tetra-
5	penta-
6	hexa-
7	hepta-
8	octa-
9	nona-
10	deca-

Examples

CO₂

→ **carbon dioxide**

N₂O₅

→ **dinitrogen pentoxide**

Key idea

Prefix = number of atoms

MOLECULAR POLARITY

A molecule's polarity depends on **two things**:

1. **Polarity of its bonds**
2. **Arrangement/geometry of those bonds**

A molecule can have polar bonds but still be nonpolar if the bond dipoles cancel.

CO₂: NONPOLAR MOLECULE

CO₂ has polar C=O bonds.

However, its shape is:

Linear

O=C=O

The two bond dipoles point in opposite directions and cancel.

Therefore:

CO₂ is nonpolar overall.

H₂O: POLAR MOLECULE

H₂O has polar O-H bonds.

Its molecular geometry is:

Bent

Because the molecule is bent, the bond dipoles do not completely cancel.

Therefore:

H₂O is polar.

Compare:

CO₂ → linear → dipoles cancel → nonpolar
H₂O → bent → dipoles don't cancel → polar

NH₃: POLAR MOLECULE

NH₃ has:

- 3 N-H bonds
- 1 lone pair on nitrogen

The lone pair affects the molecular shape.

Its molecular geometry is:

Trigonal pyramidal

The shape prevents complete cancellation of the bond dipoles.

Therefore:

NH₃ is polar.

VSEPR THEORY

VSEPR means:

Valence Shell Electron Pair Repulsion

Main idea:

Electron pairs around a central atom repel one another.

They arrange themselves as far apart as possible.

Important geometries

Electron domains	Bonding pairs	Lone pairs	Molecular geometry
2	2	0	Linear
3	3	0	Trigonal planar
4	4	0	Tetrahedral
4	3	1	Trigonal pyramidal
4	2	2	Bent

If you've read this far, very good! Now you memorize these examples ☺ Hahaha

VSEPR EXAMPLES

CO₂

2 electron domains

→ **Linear**

CH₄

4 bonding pairs

→ **Tetrahedral**

NH₃

3 bonding pairs + 1 lone pair

→ **Trigonal pyramidal**

H₂O

2 bonding pairs + 2 lone pairs

→ **Bent**

INTRAMOLECULAR VS. INTERMOLECULAR FORCES

This is another major concept in the test.

Intramolecular forces

Within a molecule

They hold atoms together.

Examples:

- Covalent bonds
- Ionic interactions within ionic compounds

For H₂O:

O—H bonds are intramolecular.

Intermolecular forces

Between molecules

They attract neighboring molecules.

Example:

Water molecule ↔ water molecule

These forces affect properties such as:

- boiling point
- melting point
- vapor pressure
- surface tension
- evaporation

Easy way to remember:

INTRA = inside

INTER = between

Tandaan niyo yan. Antagal na nian

PHYSICAL CHANGE VS. CHEMICAL CHANGE

Physical changes

The molecules remain chemically intact.

Examples:

- melting
- freezing
- evaporation
- boiling

When water boils:

H₂O(l) → H₂O(g)

The H₂O molecules remain H₂O.

The process mainly overcomes **intermolecular attractions**.

Chemical changes

Chemical bonds are broken and/or new substances form.

Example:

Hydrogen peroxide decomposing

This involves changes to the molecules themselves.

Therefore, **intramolecular bonds** are involved.

Remember:

Boiling water does NOT normally break the O—H covalent bonds.

TYPES OF INTERMOLECULAR FORCES

Three important types:

1. **London dispersion forces**
2. **Dipole-dipole forces**
3. **Hydrogen bonding**

LONDON DISPERSION FORCES

London dispersion forces result from temporary fluctuations in electron distribution. They occur in:

ALL atoms and molecules

Even nonpolar molecules experience London dispersion forces.

Example

CH₄

CH₄ is nonpolar, but CH₄ molecules can still attract each other through London dispersion forces.

Polarizability (*we haven't used this word before. Just remember the key idea below and you'll remember the rest with LDF*)

Larger electron clouds are generally more easily distorted.

Therefore:

Greater polarizability → stronger London dispersion forces

So if molecule X has a larger electron cloud than molecule Y:

X generally has stronger London dispersion forces.

DIPOLE-DIPOLE FORCES

These occur between **polar molecules**.

Polar molecules have:

- a partially positive region
- a partially negative region

These opposite partial charges attract.

Example:

HCl

H is partially positive.

Cl is partially negative.

Neighboring HCl molecules experience:

Dipole-dipole attractions

HYDROGEN BONDING

Hydrogen bonding is a particularly strong type of intermolecular attraction.

It occurs when H is bonded to a highly electronegative atom, especially:

- **N , O , F**

Common examples

- H_2O , NH_3 , HF

Water

Water molecules form hydrogen bonds with neighboring water molecules.

This contributes to water's relatively high boiling point.

STRENGTH OF INTERMOLECULAR FORCES

In general, stronger intermolecular attractions mean:

Higher boiling point

More energy is required to separate the molecules.

Lower vapor pressure

Fewer molecules escape into the gas phase.

Slower evaporation

Molecules are held more strongly.

BOILING POINT

If Liquid X has stronger intermolecular forces than Liquid Y:

X generally has:

- higher boiling point
- lower vapor pressure
- slower evaporation

Example

If two liquids have boiling points:

- 35°C
- 78°C
- 118°C

The 118°C liquid likely has the **strongest overall intermolecular attractions**, assuming other factors are reasonably comparable.

VAPOR PRESSURE

Vapor pressure is related to the tendency of molecules to escape from a liquid into the gas phase.

Stronger intermolecular forces:

→ molecules escape less easily

→ **lower vapor pressure**

Weaker intermolecular forces:

→ molecules escape more easily

→ **higher vapor pressure**

EVAPORATION

A liquid that evaporates rapidly generally has:

- weaker intermolecular attractions
- higher vapor pressure
- lower boiling point

Example

If perfume evaporates rapidly:

It likely has relatively **weak intermolecular attractions and higher vapor pressure** compared with a slower-evaporating liquid.

SURFACE TENSION

Surface tension is the tendency of liquid surfaces to resist being stretched or pulled apart.

It is caused by attractions between molecules at the liquid surface.

Example:

Water can form rounded droplets because of its relatively strong intermolecular attractions.

Practice Asking yourself with these:

Question 1: Is it ionic or covalent?

Metal + nonmetal?

→ Usually **ionic**

Nonmetal + nonmetal?

→ Usually **covalent**

Question 2: Are electrons transferred or shared?

Transferred

→ Ionic

Shared

→ Covalent

Question 3: Is the bond polar?

Compare electronegativities.

Similar → **nonpolar**

Different → **polar**

Question 4: Is the molecule polar?

Ask:

Does the molecular geometry allow the bond dipoles to cancel?

Yes → **nonpolar**

No → **polar**

Question 5: What is the molecular geometry?

Count electron domains around the central atom.

2 → **Linear**

4 domains, 3 bonds + 1 lone pair → **Trigonal pyramidal**

4 domains, 2 bonds + 2 lone pairs → **Bent**

4 bonds → **Tetrahedral**

Question 6: What intermolecular force is present?

All molecules → **London dispersion**

Polar molecules → **Dipole-dipole**

H bonded to N, O, or F → **Hydrogen bonding**

Sana masaya na kayo sa reviewer niyo. Pag to di niyo binasa ewan ko na lang talaga.