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. **CO_2 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. $CaCO_3$ has ionic attraction between **Ca^{2+} and CO_3^{2-}**, with covalent bonds within CO_3^{2-}.
8. The **octet rule** involves achieving a stable valence-shell arrangement.
9. CO_2 is written **O=C=O**.
10. NH_3 has **3 bonds + 1 lone pair**.
11. H_2O has **2 bonds + 2 lone pairs**.
12. CO_2 is **linear and nonpolar**.
13. H_2O is **bent and polar**.
14. NH_3 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. H_2O 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

$Na \rightarrow Na^+ + e^-$

Sodium loses one electron.

$Cl + e^- \rightarrow Cl^-$

Chlorine gains one electron.

Therefore:

$Na^+ + Cl^- \rightarrow NaCl$

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.

H—Cl

The electrons are shared rather than completely transferred.

Examples

- H_2O
- CO_2
- CH_4
- NH_3
- Cl_2
- O_2

Molecular compounds

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

For example:

CO_2

consists of individual CO_2 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: $CaCO_3$

Calcium carbonate contains:

Ca^{2+} and **CO_3^{2-}**

There is an **ionic attraction between Ca^{2+} and CO_3^{2-}**.

However, the atoms **inside CO_3^{2-} are connected by covalent bonds**.

So $CaCO_3$ 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 → Cl
Na loses an electron and Cl gains it.
Covalent bonding
Electron sharing
Example:
H—Cl
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:
Cl—Cl
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:
H—F
Fluorine attracts the shared electrons more strongly than hydrogen.

Bond polarity examples
Cl—Cl
Same atoms → same electronegativity
Nonpolar
C—C
Same atoms → same electronegativity
Nonpolar
C—O
Oxygen is more electronegative than carbon.
Polar
H—F
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: CH_4
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 CH_4:
C = central atom
Step 3: Connect atoms with single bonds
H—C—H
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

CH_4 — Methane
Carbon forms **four single bonds**.
Four C–H bonds
Carbon has a complete octet.

NH_3 — 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.

H_2O — Water
Oxygen has:

- 2 bonding pairs
- 2 lone pairs

Therefore:
2 O–H bonds + 2 lone pairs

CO_2 — Carbon dioxide
The correct Lewis structure is:
O=C=O
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:

- H_2
- O_2
- N_2
- F_2
- Cl_2
- Br_2
- I_2

Example: Cl_2
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**
$CaCl_2$
→ **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: $FeCl_3$
Each chloride = −1
Three chlorides:
3(−1) = −3
Therefore Fe must be:
Fe^{3+}
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:
Ca^{2+}
Nitrate:
NO_3^-
We need two nitrate ions to balance one Ca^{2+}:
$Ca(NO_3)_2$
Why parentheses?
Because there are **two complete nitrate ions**.

Example: Aluminum oxide
Aluminum:
Al^{3+}
Oxygen:
O^{2-}
The smallest combination that gives zero total charge is:
2 Al^{3+} = +6
3 O^{2-} = −6
Therefore:
Al_2O_3

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

Examples
CO_2
→ **carbon dioxide**
N_2O_5
→ **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.

CO_2: NONPOLAR MOLECULE

CO_2 has polar C=O bonds.
However, its shape is:
Linear
O=C=O
The two bond dipoles point in opposite directions and cancel.
Therefore:
CO_2 is nonpolar overall.

H_2O: POLAR MOLECULE

H_2O has polar O–H bonds.
Its molecular geometry is:
Bent
Because the molecule is bent, the bond dipoles do not completely cancel.
Therefore:
H_2O is polar.
Compare:

CO_2 → linear → dipoles cancel → nonpolar
H_2O → bent → dipoles don't cancel → polar

NH_3: POLAR MOLECULE

NH_3 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:
NH_3 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

CO_2
2 electron domains
→ **Linear**
CH_4
4 bonding pairs
→ **Tetrahedral**
NH_3
3 bonding pairs + 1 lone pair
→ **Trigonal pyramidal**
H_2O
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 H_2O:
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:
$H_2O(l)$ → $H_2O(g)$
The H_2O molecules remain H_2O.
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
CH_4
CH_4 is nonpolar, but CH_4 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.