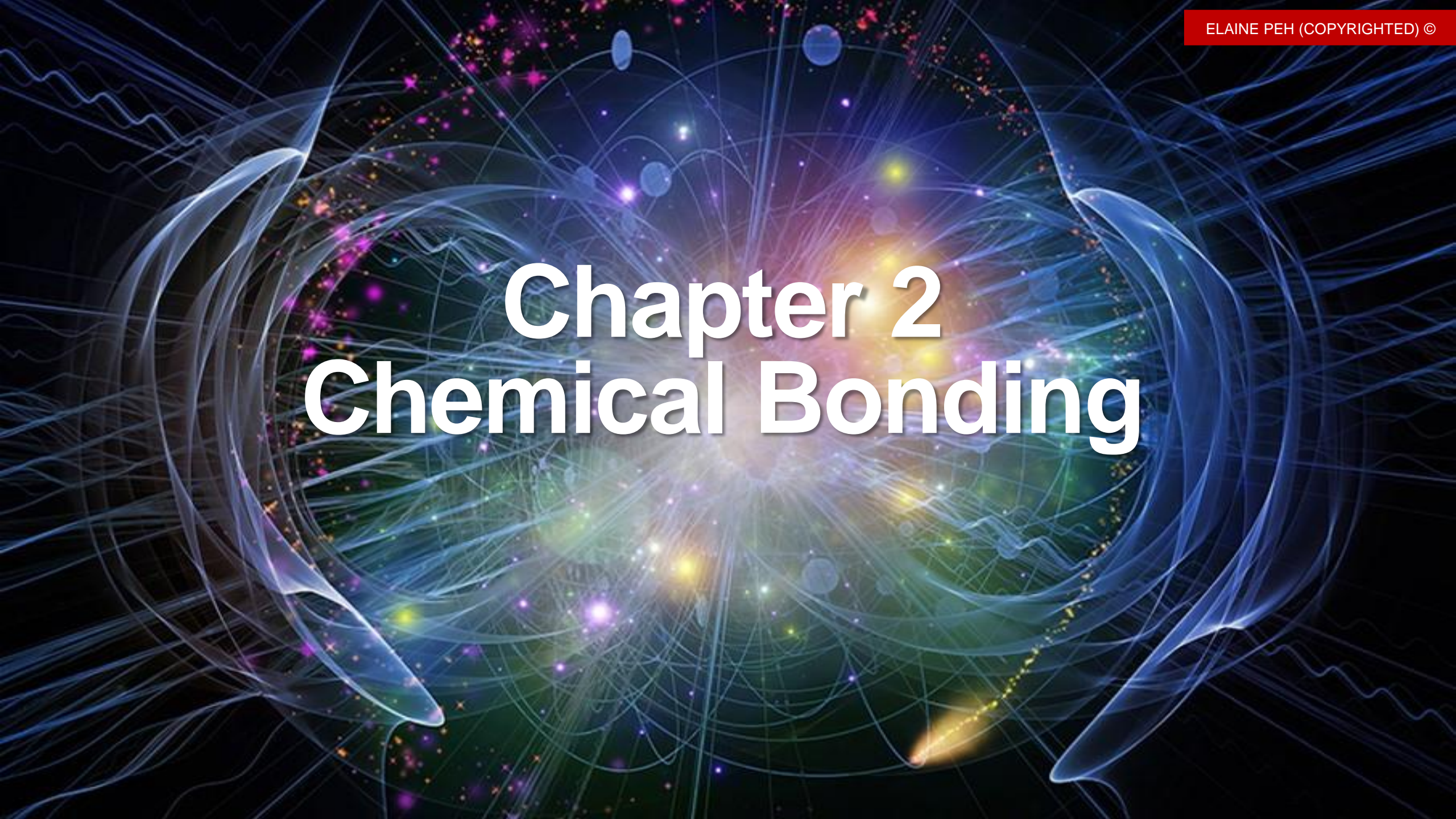


# Chapter 2

# Chemical Bonding



# CHAPTER ANALYSIS



FOCUS

- Content Heavy Chapter
- 5 **key** concepts



EXAM

- Always tested
- Standard template applies when answering theory-based questions



WEIGHTAGE

- Heavy overall weightage
- Lots of practice require for drawing of dot-and-cross diagrams



KEY CONCEPT

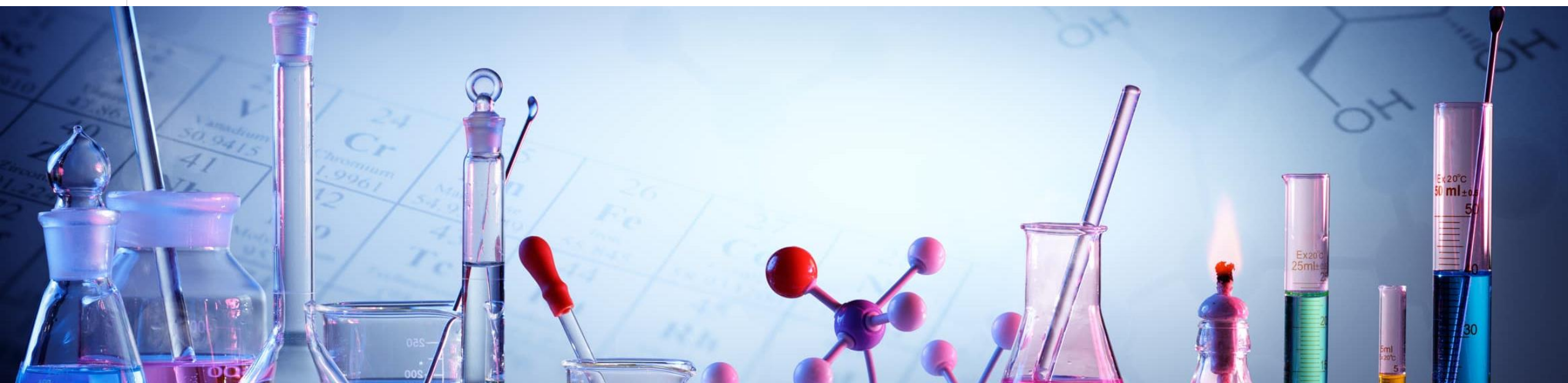
# Intramolecular Bonding

## Intermolecular Forces of Attraction

### Physical Properties of Compounds based on Structure & Bonding

### Dot-and-Cross Diagrams

### Shapes and Bond Angles of Molecules



# Intramolecular Bonding



Intramolecular Bonding are electrostatic forces of attraction form **within** the molecules, **between** individual atoms.

## METALS

Metals have **giant metallic lattice structure** consisting of metal cations that are arranged in a lattice and surrounded by mobile '**sea**' of **delocalised electrons**.

## METALLIC BONDS

It refers to the **electrostatic forces of attraction** between a lattice of **positive (metal) ions** and the '**sea**' of **delocalised electrons**.

## IONIC COMPOUNDS

Ionic compound is formed **between a metal cation and a non-metal anion**, whereby the **valence electrons from the metallic atoms** are **completely transferred** to the non-metallic atom.

## IONIC BONDS

It refers to the **electrostatic forces of attraction** between the **oppositely charged ions (cations and anions)**.

# Intramolecular Bonding



## COVALENT COMPOUNDS

Covalent compound consists of **non-metals atoms** (Note: There are exceptions cases;  $AlCl_3$  and  $BeCl_2$ ) bonded together by **covalent bonds**.

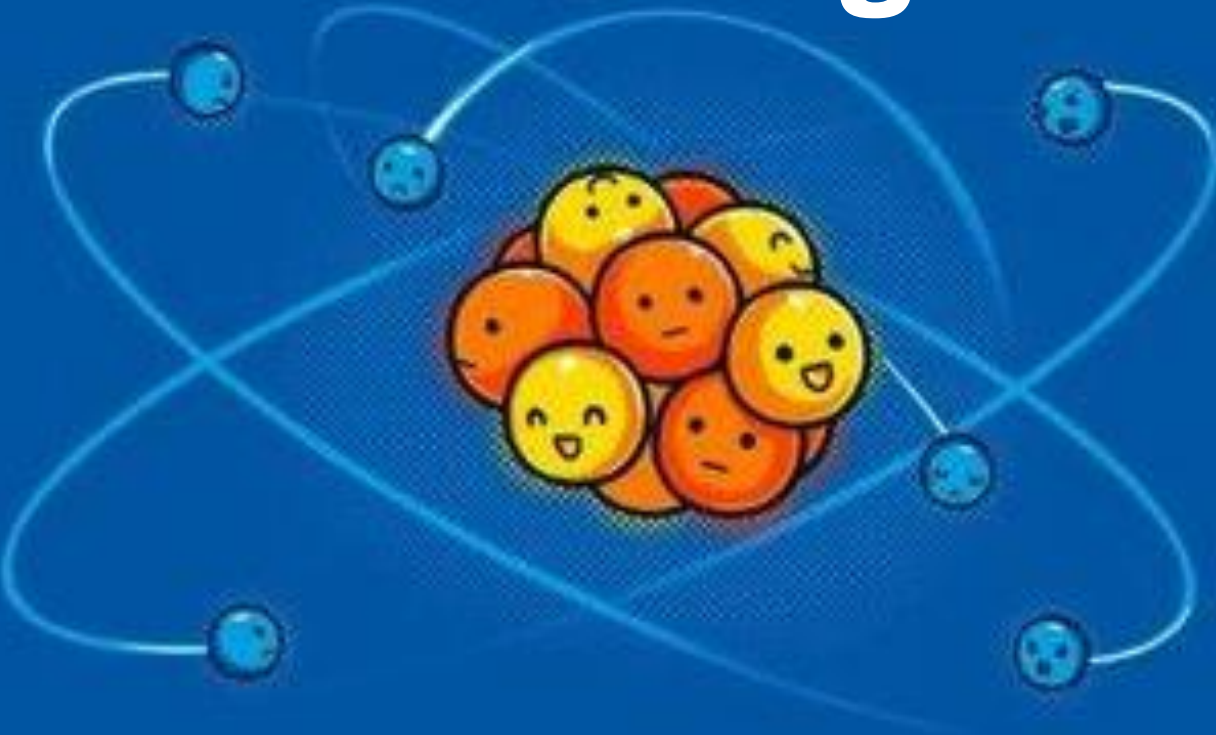
## COVALENT BONDS

It refers to the **electrostatic forces of attraction** between the **shared pair of electrons** and the **positively charge nuclei**.

### Types of Covalent Bonds

1. Both singly occupied with an electron (normal covalent bond)
  - i. **Sigma bond ( $\sigma$  bond)**: formed when two orbitals overlap head-on overlap
  - ii. **Pi bond ( $\pi$  bond)**: formed only when the p orbitals overlap sideways
2. One is filled with a pair of electrons while the other is an empty orbital (dative covalent bond)
  - It occurs when the **shared pair of electron** is **provided by only one of the bonding atom**
  - The atom must have a **lone pair of electron for donation** while the other atom must have an **empty orbital available** to accept the pair of electron

# Intramolecular Bonding



## Types of Covalent Bonds

### 1. Non-Polar Covalent Bonds

- Occur when **two atoms of the same electronegativity** (the tendency to attract electrons) form a covalent bond, there is an **equal sharing of the bonding electrons** between the two nuclei.

### 2. Polar Covalent Bonds

- **Occur** when **two atoms of different electronegativity** form a covalent bond, the **bonding electrons is not equally shared** and the **bond becomes polar** and is said to **possess ionic characteristic** whereby the **atom of greater electronegativity** now **carries a  $\delta^-$  charge** while the **atom of lower electronegativity carries a  $\delta^+$  charge**.
- The **degree of polarity** of the bond its **measure by its dipole moment**  $\rightarrow$ , with the arrow pointing towards the more electronegative atom



# Intramolecular Bonding



## Factors Affecting the Bond Strength

### 1. Bond Order

- **Bond order** refers to the **number of covalent bonds** between 2 bonded atoms
- **Multiple bonds stronger than single bonds** as more electrons shared hence stronger electrostatic forces of attraction

### 2. Bond Length

- **Bond length** is the **distance between the nuclei of the two atoms** involved in the covalent bond.
  - **Smaller atoms form shorter bond length** as overlap between valence orbitals are more effective
- Hence **stronger electrostatic forces of attraction** from nucleus resulting in **stronger bond strength**

# Intramolecular Bonding



## Types of Covalent Compounds

### 1) Simple Covalent Compound

Molecules held together by **strong electrostatic attraction between the atoms** (intramolecular forces of attraction) but have **weak electrostatic attraction between the molecules** (intermolecular forces of attraction).

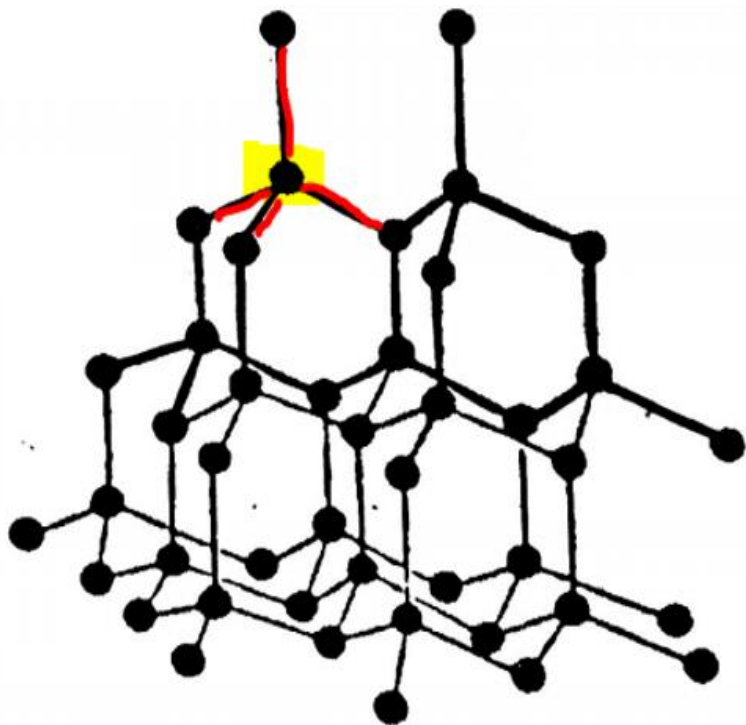
### 2) Giant Covalent Compound

Molecules held together by **strong covalent bonds** (intramolecular forces of attraction) between the atoms in a 3D network. (Eg: **Diamond & Graphite**)



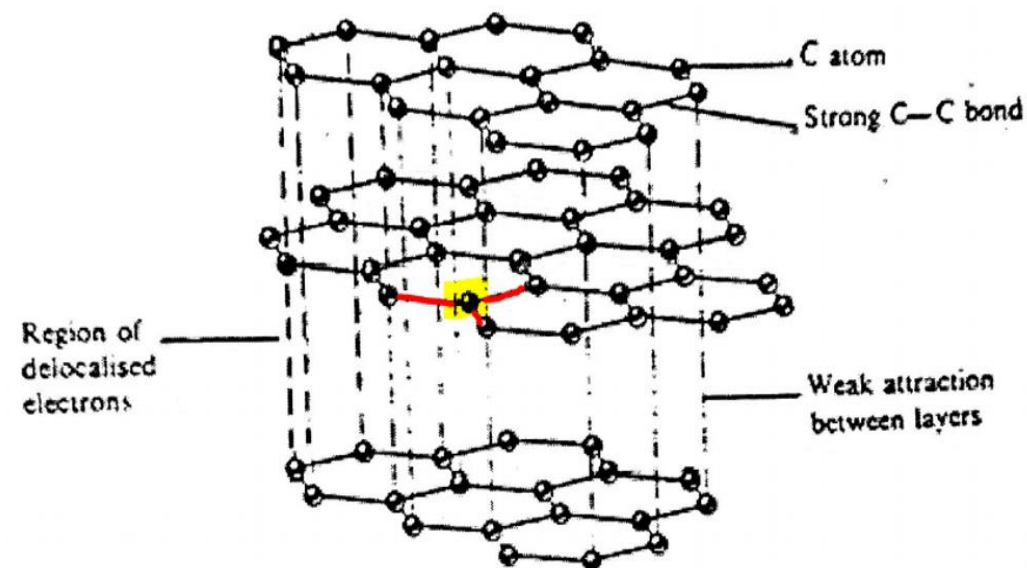
## DIAMOND

Diamond has a giant covalent molecular structure with the **C atoms held together by strong electrostatic attraction between one another**. Each C atom is covalently bonded to 4 other C atoms in a 3D tetrahedral arrangement.

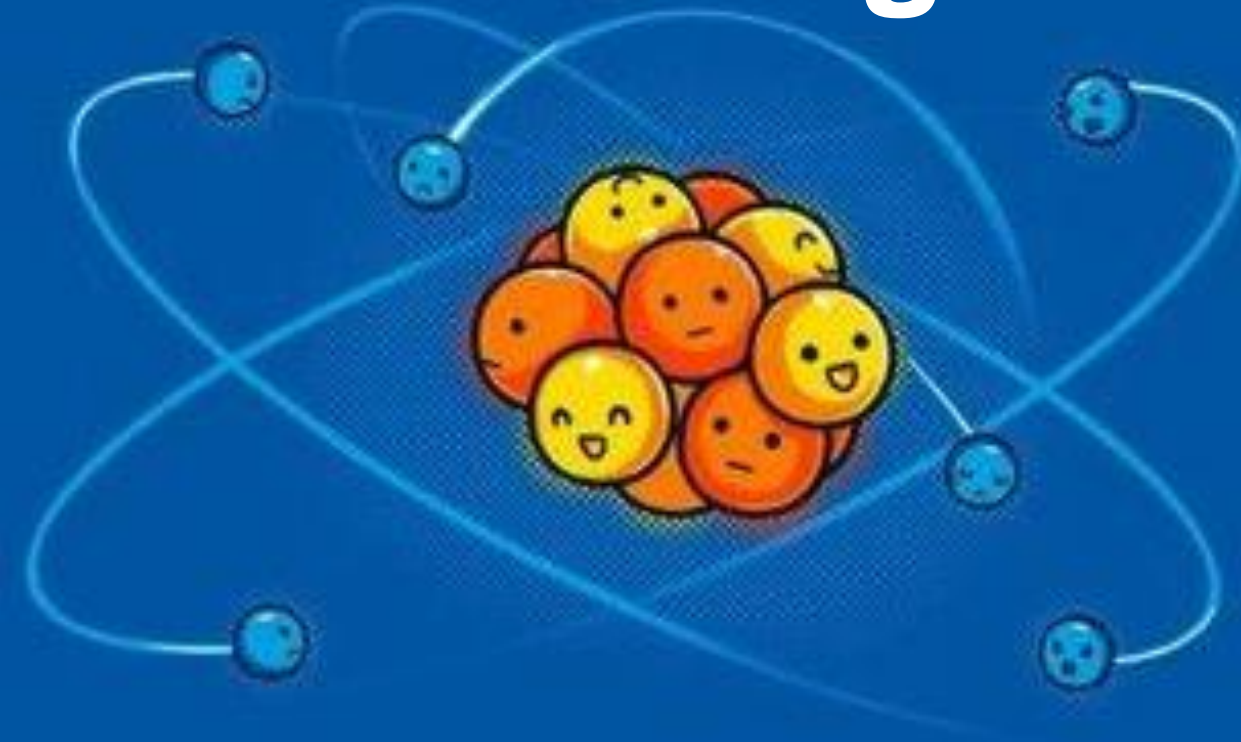


## GRAPHITE

Graphite has a giant covalent molecular structure with the **C atoms held together by strong electrostatic attraction between one another** arranged in layers of interconnected hexagonal rings. **Within each layer, each C atom is covalently bonded to 3 other atoms**, using only 3 out of 4 of its valence electrons. The **4th electron is delocalised** throughout the layer and is available for bonding **between adjacent layer**, which are held together by weaker intermolecular forces.

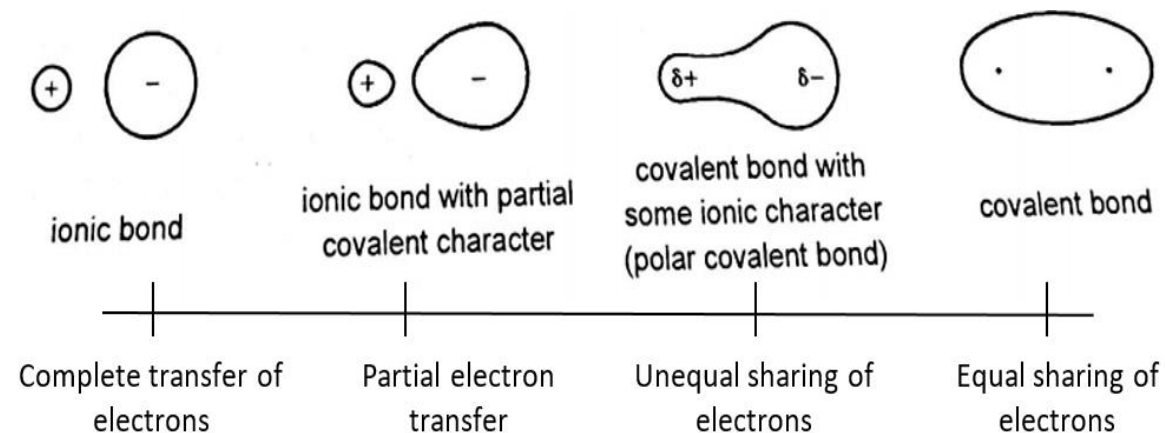


# Intramolecular Bonding



## INTERMEDIATE BOND TYPES

The **type of bond** is determined by **difference in electronegativity between two atoms**



## Ionic Bond with Partial Covalent Characteristic

Due to their **oppositely charged nature**, the **attraction of the valence shell electrons of the anion to the cation** result in **distortion (polarisation) of the electron cloud of the anion**. This results in **partial electron sharing** resulting in the covalent characteristic.

### The degree of covalent characteristic depends on:

#### 1. Polarising Power of the Cations

- Dependent on its **charge density**  $\propto \frac{\text{charge}}{\text{size}}$

#### 2. Polarisability of the Anion

- Dependent **more on its size than charge**

## Covalent Bond with Partial Ionic Characteristic

- Referred to as **polar covalent compounds**
- They have polar covalent bonds **AND** the dipole moment from the polar covalent bonds does not cancel out (net dipole moment).
- This occur when the **polar bonds are arranged in a non-symmetrical arrangement**, thus **unable to cancel out one another**.



Intermolecular forces of attraction are electrostatic in nature and are form **between molecules**. They are **generally weaker** than intramolecular bonds formed within the molecules.

#### INSTANTANEOUS DIPOLE – INDUCE DIPOLE (id-id) INTERACTIONS

Exist for **all** covalent molecules.

#### PERMANENT DIPOLE – PERMANENT DIPOLE (pd-pd) INTERACTIONS

Exist for **only** polar covalent molecules.


#### HYDROGEN BONDING (H-BOND)

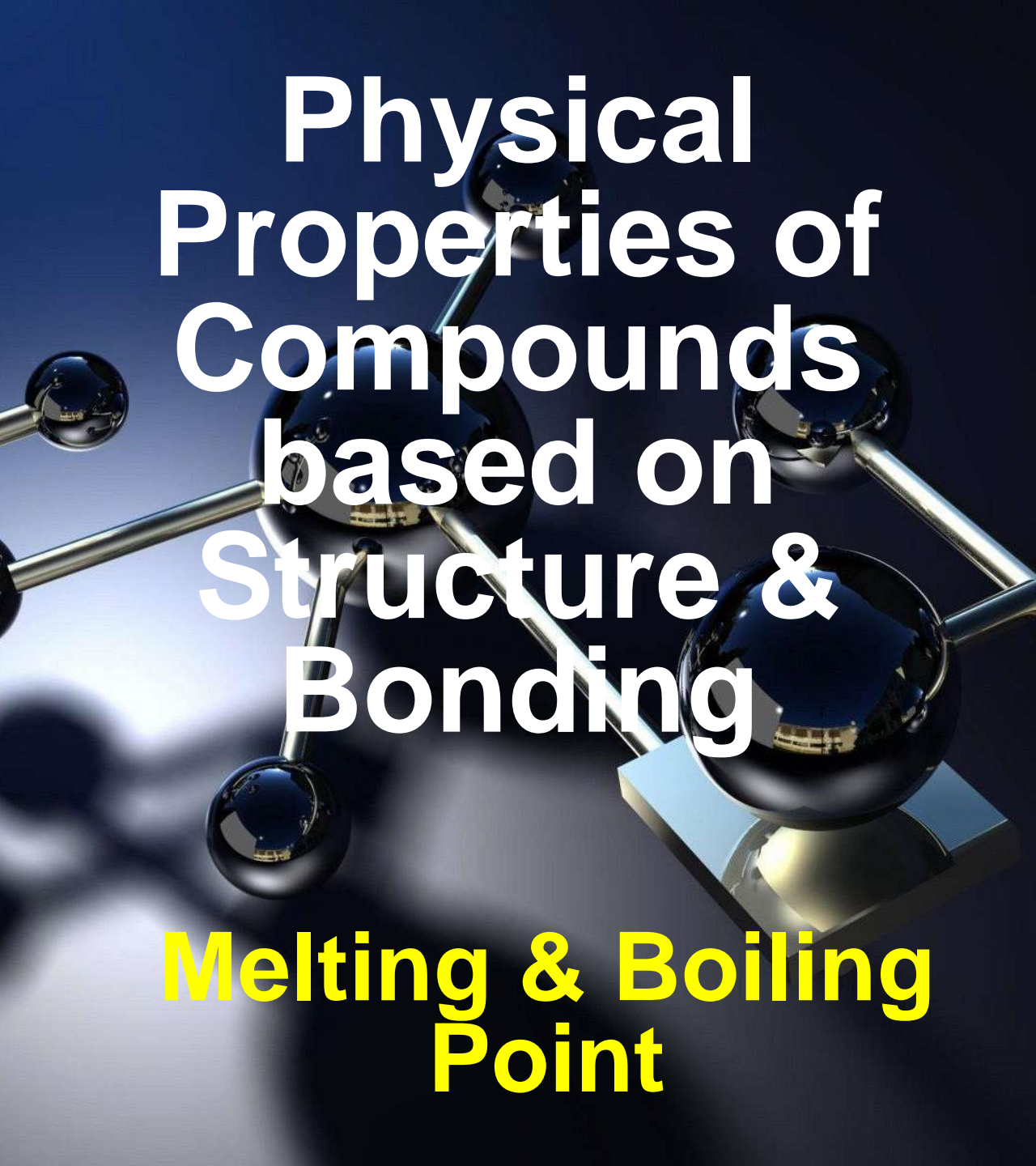
H-Bond is a **type of pd-pd interaction** that exist for only polar covalent molecules that **fulfill the following criteria**:

- The molecule must have a **H atom bonded directly to F/O/N atom**
- The **F/O/N atom** must have **at least one lone pair of electron**

The formation of **intermolecular H-Bond between molecules** result in **dimerisation of compounds** (eg: carboxylic acid).

# Intermolecular Forces of Attraction





# Physical Properties of Compounds based on Structure & Bonding

## Melting & Boiling Point

**Melting & Boiling Point** of a substance is the temperature that **measure the amount of energy required to overcome the bonds BETWEEN the particles** to break down the regular arrangement of atoms.

**Steps to Compare Melting/Boiling Point of Molecules:**

**Step 1:** Identify the structure and type of attractions present in the molecules

**Step 2:** Compare the difference in strength of attractions present

**Step 3:** State the degree of energy required (large/small amount) to take in to overcome the forces of attraction and hence link to the melting/boiling point.

# Melting & Boiling Point – Comparing between **DIFFERENT** Structures

1. **Metallic Compounds:** Consist of metallic bonds. **Usually have high melting & boiling point** as compared to other structures as **large amount of energy** is required to overcome the strong electrostatic attraction between the metal cations and 'sea' of delocalised electrons
2. **Ionic Compounds:** Consist of ionic bonds. **Usually have high melting & boiling point** as compared to other structures as **large amount of energy** is required to overcome the strong electrostatic attraction between the oppositely charged ions in the lattice structure
3. **Giant Covalent Compounds:** Consist of covalent bonds. Usually have **high melting & boiling point** as compared to other structures as **large amount of energy** is required to overcome the strong covalent bonds between the atoms in the giant lattice structure
4. **Simple Covalent Compounds:** Consist either of **id-id**, **pd-pd** or/and **H-bond**. Usually have **low melting & boiling point** as compared to other structures as **small amount of energy** is required to overcome the weak electrostatic attraction between the molecules



# Melting & Boiling Point – Comparing between **SIMILAR** Structures

## Metallic Compounds

### 1. Number of delocalised electrons

- The **more delocalised electrons** a metal have, the **stronger their electrostatic attraction** to the metal cation and thus the **stronger the metallic bond**.

### 2. Charge density of metal cation

- Charge density  $\propto$

## Ionic Compounds

### 1. Charge of the Ions

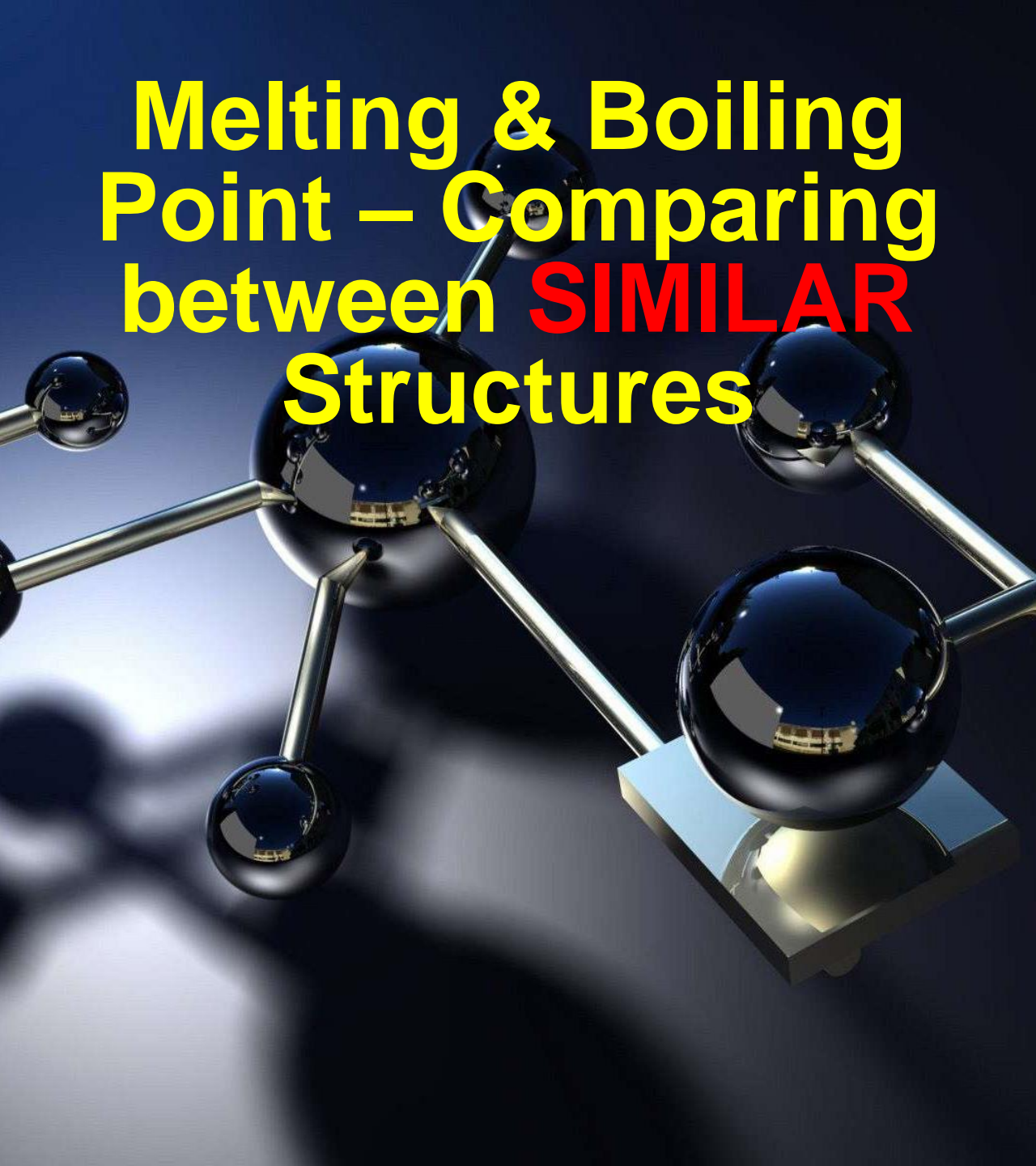
- The **higher the charge** of the ions, the **stronger the electrostatic attraction** to each other and hence the stronger the ionic bond

### 2. Size of the Ions

- The **smaller the size** of the ions, the **shorter the distance** is between each other and hence resulting in a **stronger electrostatic attraction** between the oppositely charged ions and thus a stronger ionic bond

*Note: The charge factor is a more important factor than the size factor*

# Melting & Boiling Point – Comparing between **SIMILAR** Structures



## Simple Covalent Compounds

- Simple Covalent Compounds consist either **H-bond**, **pd-pd** or/and **id-id**
- The strength of bondings between molecules decrease in the following order → H-Bond > pd-pd > id-id

### H-Bond

1. The magnitude of the dipole moment of the H-(F/N/O) bond
  - The **greater the magnitude** of the dipole moment, the **stronger the H-Bond** is between molecules
  - **Dipole moment of the H-(F/N/O) bond:**  $\text{H-F} > \text{H-O} > \text{H-N}$

*Recap: The larger the difference in electronegativity of the bonding atoms, the greater the magnitude of the dipole moment*

2. The number of H-Bond formed
  - The **more intermolecular H-Bonds** each molecule is able to form between one another, the **stronger the H-Bonds** are
3. Formation of intramolecular H-Bond within the molecules
  - The **formation of intramolecular H-Bonds decrease the availability of the H-atom** and the lone pairs available for the formation of intermolecular H-Bonds.
  - This thus **decrease the number of intermolecular H-Bonds formed**, hence decreasing the strength of the H-bonding between molecules.

# Melting & Boiling Point – Comparing between **SIMILAR** Structures

## Simple Covalent Compounds

### pd-pd interactions

#### 1. The magnitude of the molecule's net dipole moment

- The greater the magnitude of the dipole moment, the more polar the molecule is. Hence the stronger the pd-pd interaction is between the molecules.

*Recap: The larger the difference in electronegativity of the bonding atoms, the greater the magnitude of the dipole moment and the more polar the bond is*

## Simple Covalent Compounds

### id-id interactions

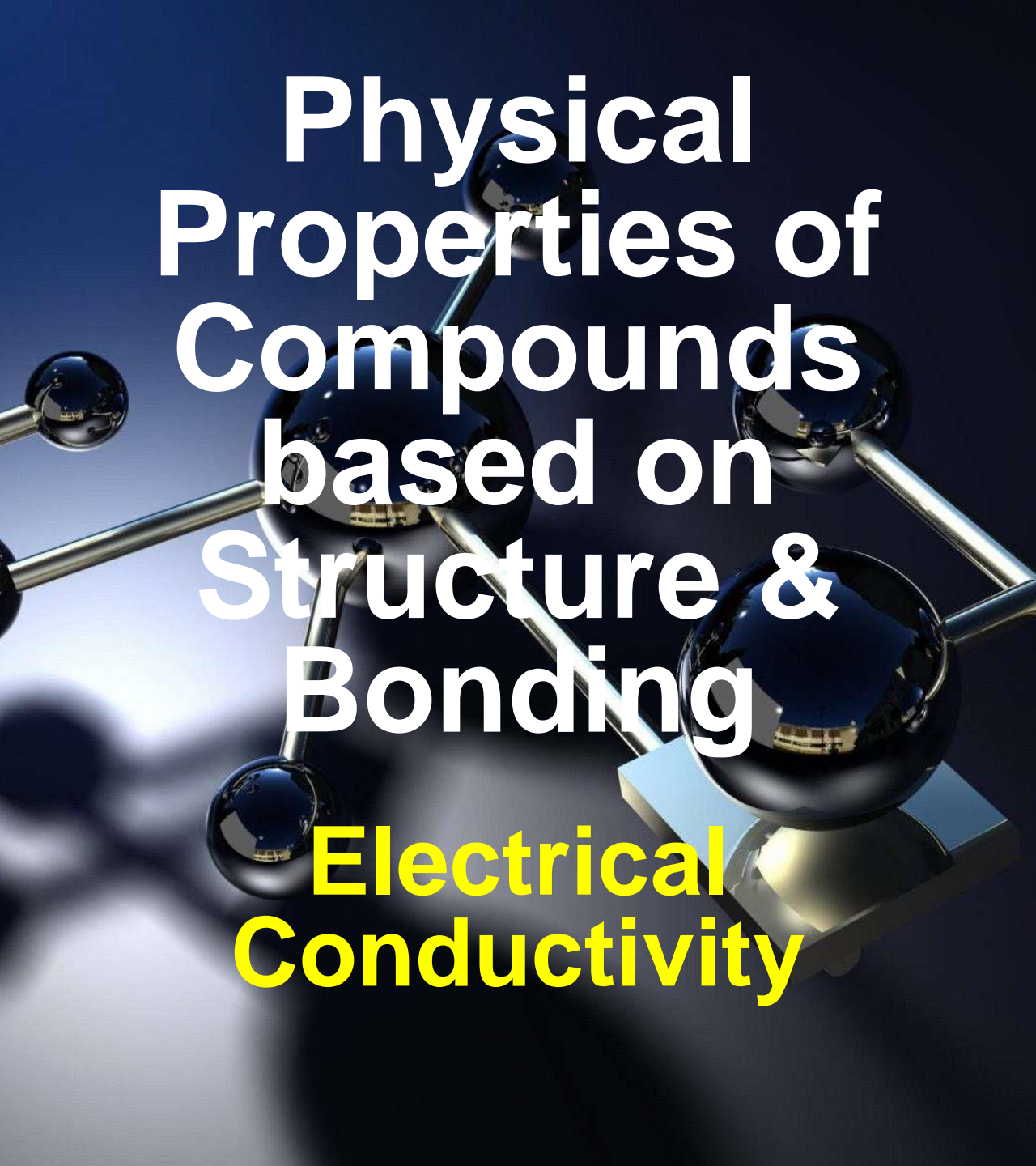
#### 1. Size of Electron Cloud Surface Area for Contact

- The **bigger the size** of the electron cloud, the easier the electron cloud **can be polarised** resulting in **stronger the id-id interactions** formed between molecules.

#### 2. Surface Area for Contact

- The **greater the surface area** the molecules have in contact between one another, the **more extensive the id-id interactions can be formed** resulting in **stronger the id-id interactions** formed between molecules.





# Physical Properties of Compounds based on Structure & Bonding

## Electrical Conductivity

### 1. Metallic Compounds

Electrical Conductor in both Solid & Molten State

- Consist of metallic bonds
- Mobile delocalised electrons act as charge carriers
- Conductivity increases as ionisable valence shell / delocalised electrons increases (more mobile charged carriers)
- Mobile delocalised electrons also allow electron diffusion, hence thermal energy can be picked up and transferred

### 2. Ionic Compounds

Electrical Conductor only in Molten & Aqueous State but not in Solid State

- Consist of ionic bonds
- **Molten & Aqueous State:** the oppositely charged ions are free to move about and hence act as charge carriers to conduct electricity
- **Solid State:** the oppositely charged ions are held at fixed positions and hence unable to act as mobile charge carrier

# Physical Properties of Compounds based on Structure & Bonding

## Electrical Conductivity

### 3. Giant Covalent Compounds

#### a. Diamond

##### **Non-conductor of electricity**

- As all valence electrons are used to form covalent bonds, they are unable to move freely and hence unable to act as mobile charge carrier.

#### b. Graphite

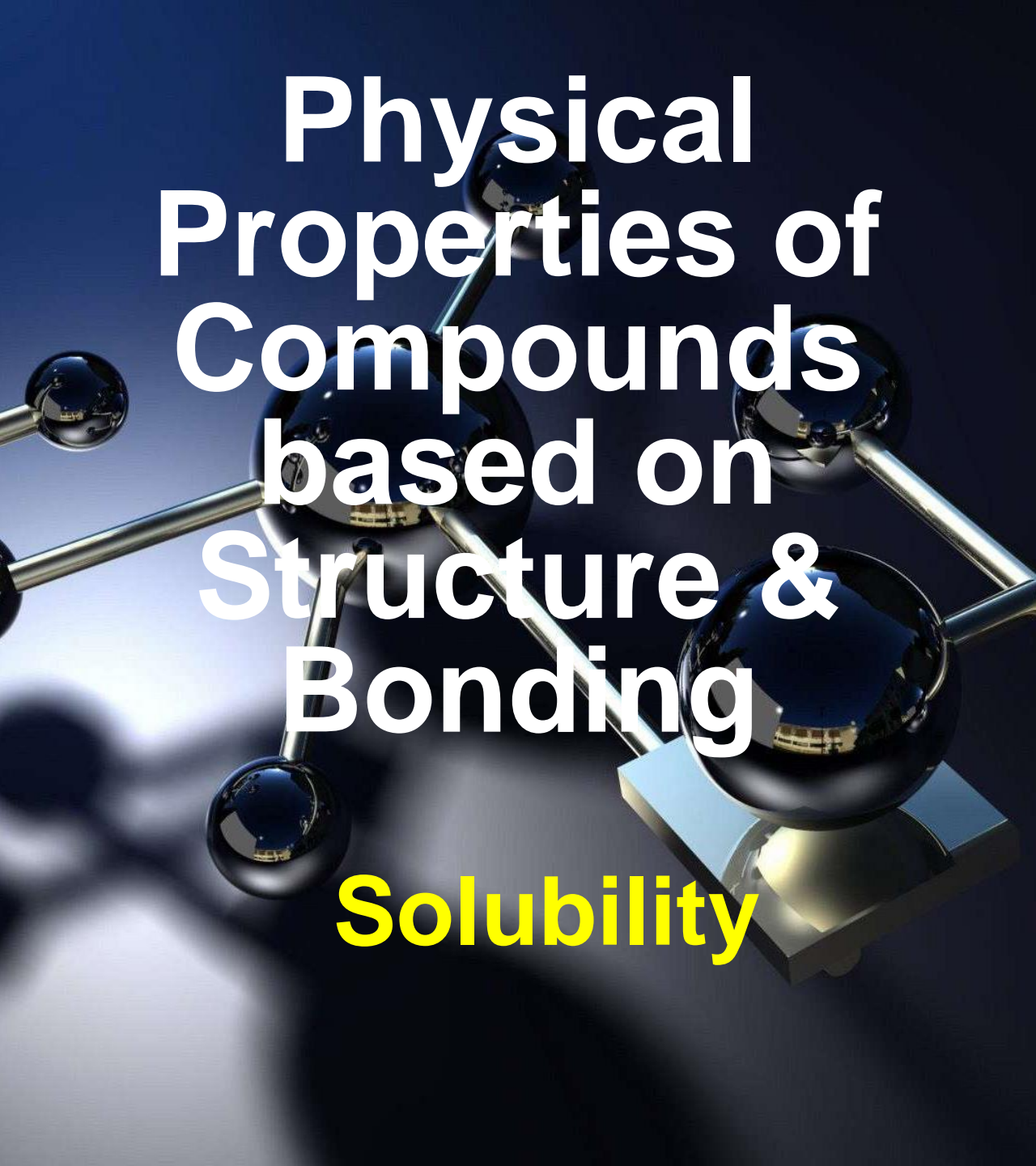
##### **Good conductor of electricity**

- The presence of delocalised electrons (the 4<sup>th</sup> electrons) between the layers act as mobile charge carrier. Hence graphite sheets can conduct electricity within the layers. However, it does not conduct electricity across layers (right angles to each other)

### 4. Simple Covalent Compounds

##### **Non-conductor of electricity**

- All the electrons are used up for bonding hence there are no mobile charge carriers available.



# Physical Properties of Compounds based on Structure & Bonding

## Solubility

### DETERMINE SOLUBILITY

To determine if the solute is soluble in a solvent, we need to first identify the:

- **Solute-solute interaction**
- **Solvent-solvent interaction**
- **Solute-solvent interaction**

#### **It is soluble only when**

- Energy given out during Solute-solvent interaction  $\geq$  Energy taking in during Solvent-solvent interaction
- Energy given out during Solute-solvent interaction  $\geq$  Energy taking in during Solute-solute interaction



# Physical Properties of Compounds based on Structure & Bonding

## Solubility

### 1. Ionic Compounds

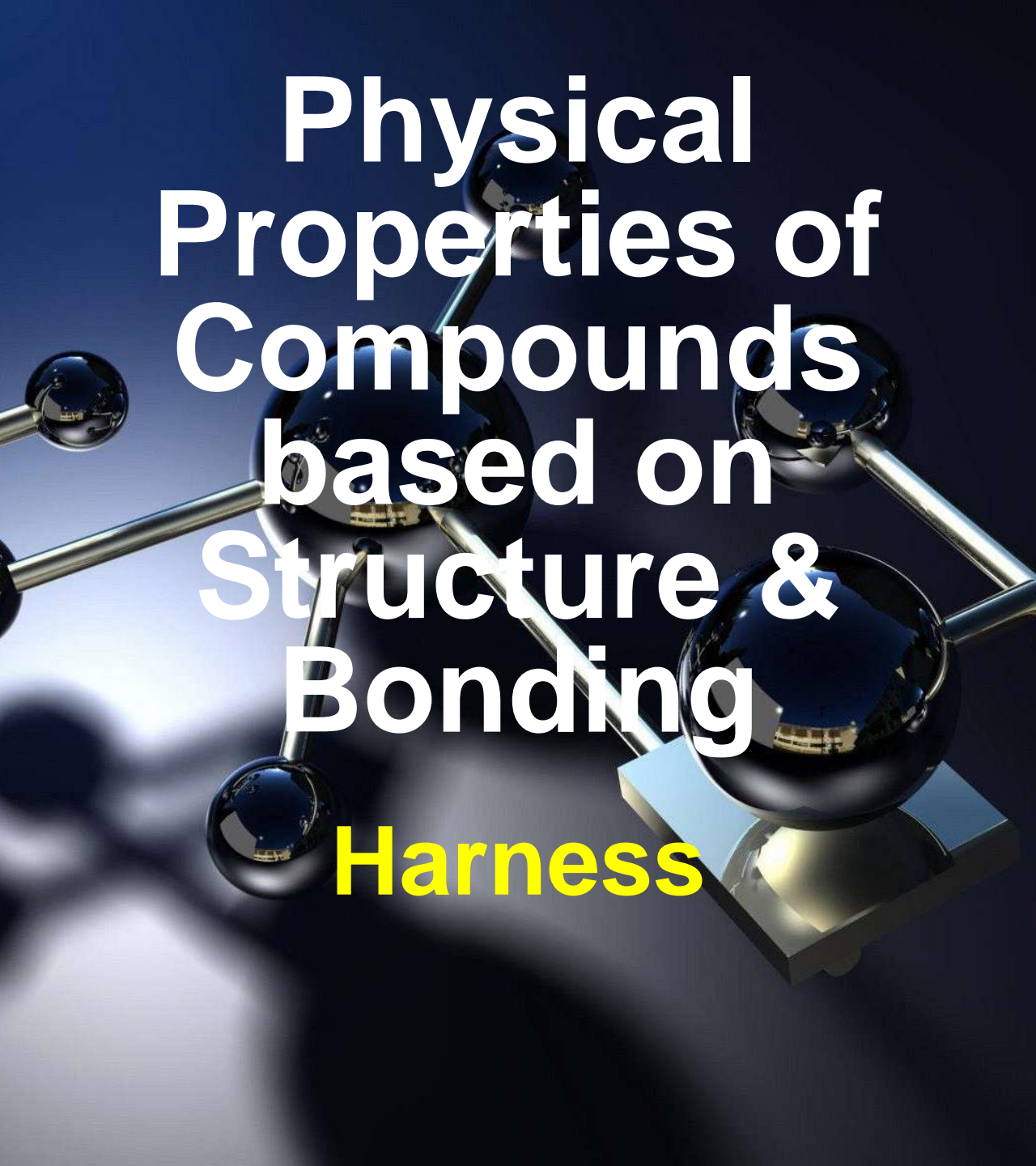
- **Soluble in polar solvents (water)**
  - Solute-solvent interaction: ionic bond
  - Solute-solute interaction: ionic bond
  - Solvent-solvent interaction: H-bond
- **Insoluble in non-polar solvents**
  - Solute-solvent interaction: pd-pd/id-id interaction
  - Solute-solute interaction: ionic bond
  - Solvent-solvent interaction: pd-pd/id-id interaction

### 2. Giant Covalent Compounds

- a. Diamond
  - Insoluble in polar solvents (water)**
    - Solute-solvent interaction: pd-pd/id-id interaction
    - Solute-solute interaction: covalent bonding
    - Solvent-solvent interaction: H-bond
- b. Graphite
  - Insoluble in polar solvents (water)**
    - Solute-solvent interaction: pd-pd/id-id interaction
    - Solute-solute interaction: covalent bonding
    - Solvent-solvent interaction: H-bond

### 3. Simple Covalent Compounds

- **Insoluble in polar solvents (water)**
  - Solute-solvent interaction: pd-pd/id-id interaction
  - Solute-solute interaction: pd-pd/id-id interaction
  - Solvent-solvent interaction: H-bond
- **Soluble in non-polar solvents**
  - Solute-solvent interaction: pd-pd/id-id interaction
  - Solute-solute interaction: pd-pd/id-id interaction
  - Solvent-solvent interaction: pd-pd/id-id interaction



# Physical Properties of Compounds based on Structure & Bonding

## Harness

### 1. Metallic Compounds

#### **Hard, but not Malleable & Ductile (not brittle)**

Hard: Metals do not deform easily under physical stress/force due to their strong metallic bonds

Malleable & Ductile: The 'sea' of delocalised electrons prevent repulsion (between cations) and allow the formation of new bonds between the cation and electrons when layers of metal cations slide past each other. This enable the metallic bonding to remains intact even as a force is applied.

### 2. Ionic Compounds

#### **Hard & Brittle**

Hard: Ionic compound do not deform easily under physical stress/force due to their strong ionic bonds

Brittle: As the ions are arranged in an orderly lattice structure, ions of similar charges might come together during the displacement of the layers of ions when a forced is applied. Repulsion between the like-charges will thus result the two portions of the lattice structure to fall apart.

# Physical Properties of Compounds based on Structure & Bonding

## Harness

### 3. Giant Covalent Compounds

#### a. Diamond

##### **Very Hard**

- Diamond is strong and rigid due to the strong C-C bonds formed in the 3D network.

#### b. Graphite

##### **Soft & Slippery**

- The weak intermolecular forces of attraction between the layers allow the layers to slide past one another when a force is applied.

### 4. Simple Covalent Compounds

#### **Soft & Slippery**

- The weak intermolecular forces of attraction between the layers allow the layers to slide past one another when a force is applied.



**Octet Rule** is use as a guiding principle in drawing dot-and-cross diagrams whereby the **octet structure of 8 valence electrons** is achieved whenever possible

### Ionic Compounds

**Step 1:** Identify the cation and anion and their respectively charges

**Step 2:** Use an empty shell to represent the loss of electrons in the cation and bracket it.

**Step 3:** Write its charges outside the brackets.

**Step 4:** Identify the number of valence electrons of the anion and draw it.

- The number of valence electrons corresponds to the group of the element in the Periodic Table.
- *Note: we will usually draw the electrons in a NSEW direction for the first 4 electrons*

**Step 5:** Draw the newly added electrons from the cation with another symbol and write its charges outside the brackets for the anion.

**Step 6:** Draw both ions together

# Dot-and-Cross Diagrams

## Covalent Compounds

**Step 1:** Identify the central atom

- The central atom is usually the element present in lower number **OR** has the lowest electronegativity

**Step 2:** Determine the number of valence electrons of each atom

The number of valence electrons corresponds to the group of the element in the Periodic Table.

**Step 3:** Determine the number of bonds (hence the number of electrons) each atom needs to form in order to achieve octet structure (valence electron of 8)

**Step 4:** Draw out the valence electrons including those not involved in bonding.

- Bond Pair: the electron pair that is shared between the bonded atom
- Lone Pair: the electron pair that is not shared between the bonded atom
- *Note: Alternative symbol is use to represent electrons from different element.*

**Step 5:** Check that the central atom and surrounding atoms achieve octet structure.

- If octet structure cannot be achieved by normal covalent bond, consider dative bonding.
- *Recap: There are two types of covalent bonds, (1) normal covalent bond (both singly occupied with an electron); (2) dative covalent bond (one is filled with a pair of electrons while the other is an empty orbital)*
- *Note: Dative bonding is represented by  $\rightarrow$  from the donor atom to the acceptor atom or by both electrons*



# Dot-and-Cross Diagrams



## Polyatomic Ions

### Step 1:

- For cations, electrons are generally lost from the less electronegative atom
- For anions, electrons are generally gained by the more electronegative atom

**Step 2:** Follow the steps as per drawing a covalent compound

**Step 3:** The overall charge on the ion must be included and written outside the brackets

# Dot-and-Cross Diagrams

## EXCEPTION TO THE OCTET RULE

### 1. Electron Deficient Species

- Usually occur when the central atom have less than 8 electrons in the valence shell but yet no valence electrons are available to be taken in from the neighboring atom
- Examples:  $\text{BeCl}_2$ ,  $\text{BF}_3$ ,  $\text{AlCl}_3$

### 2. Radicals

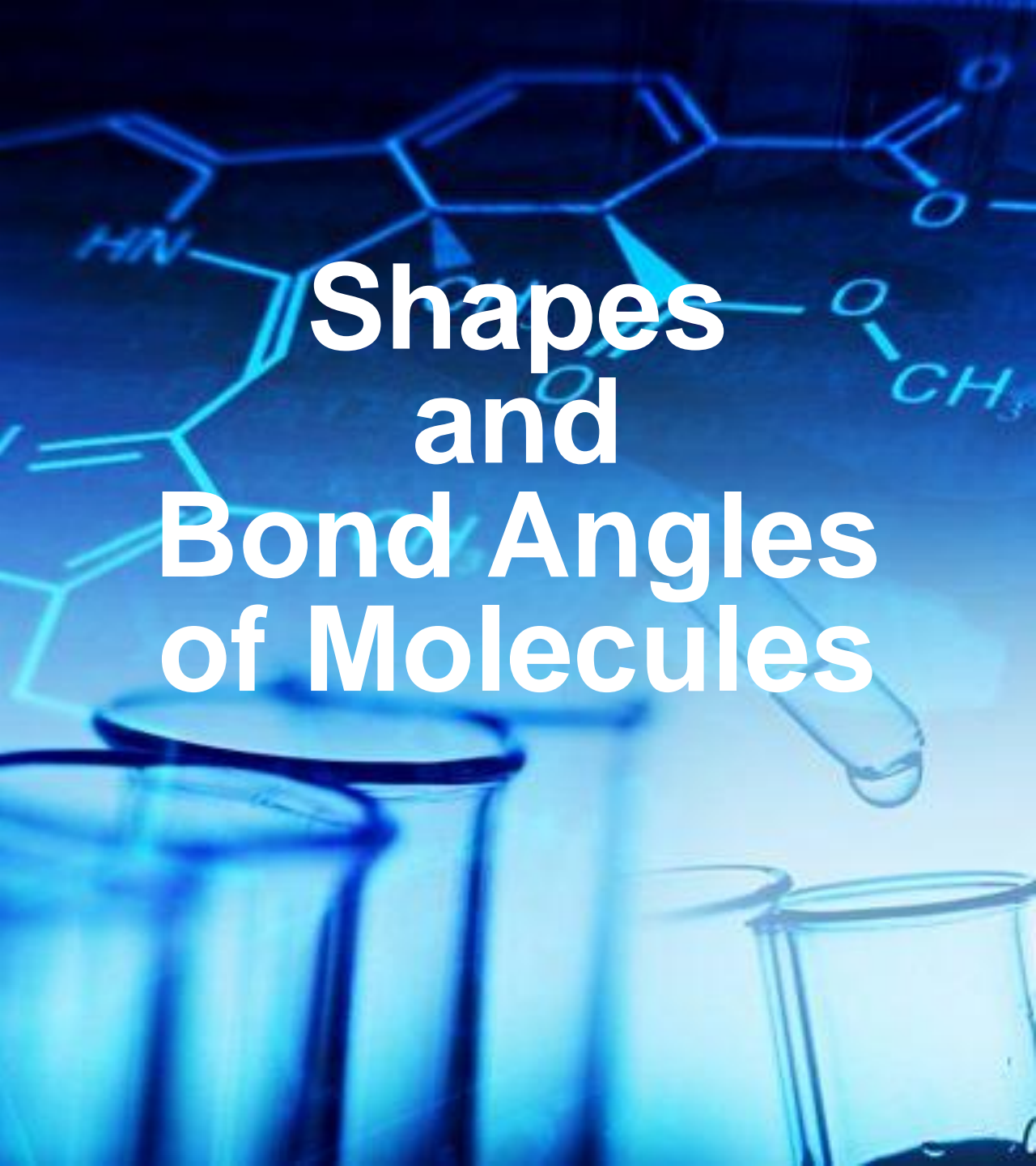
- Usually occur with atoms of odd number electrons, forming one unpaired non-bonding electron.
- Examples:  $\text{NO}$ ,  $\text{NO}_2$

### 3. Expanded Octet Structure

- Only occur for elements in Period 3 and above as such element can also use 3d subshells to form covalent bonds. Therefore they are able to expand its octet structure to take in more than 8 valence electrons (usually for the central atom).
- Examples:  $\text{PCl}_5$

# Dot-and-Cross Diagrams



The background of the slide features a blue-tinted image. In the upper portion, there are faint, glowing chemical structures, including what appears to be a complex organic molecule with a ring and a carboxylate group, and a smaller molecule with an amine group (HN) and a methoxy group (O-CH3). In the lower portion, there are several pieces of laboratory glassware, including Erlenmeyer flasks and a graduated cylinder, some containing liquids. The overall theme is chemistry.

# Shapes and Bond Angles of Molecules

## VALENCE SHELL ELECTRON PAIR REPULSION (VSEPR) THEORY

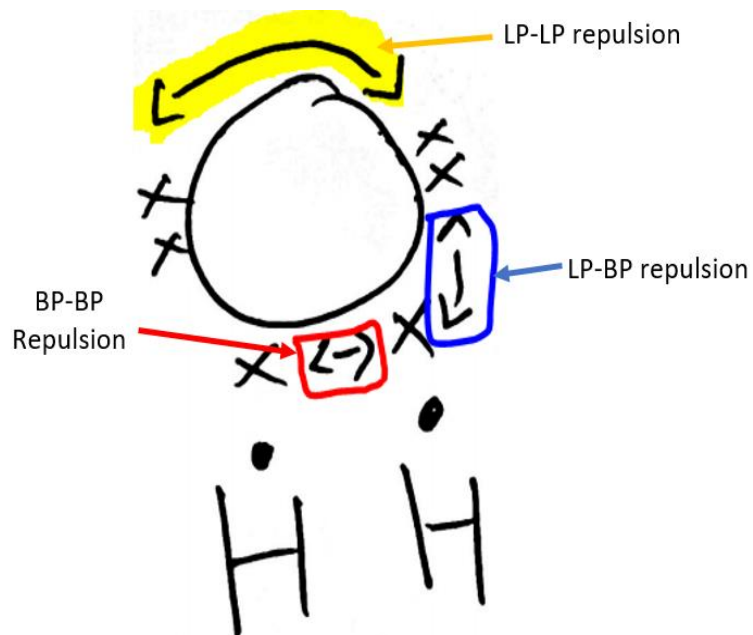
**Principle 1:** Electrons pairs arrange themselves around the central atom of a molecule as far apart as possible to minimise mutual repulsion

**Principle 2:** Repulsion between electron pairs decrease in the following order  $\rightarrow$  LP-LP > BP-LP > BP-BP

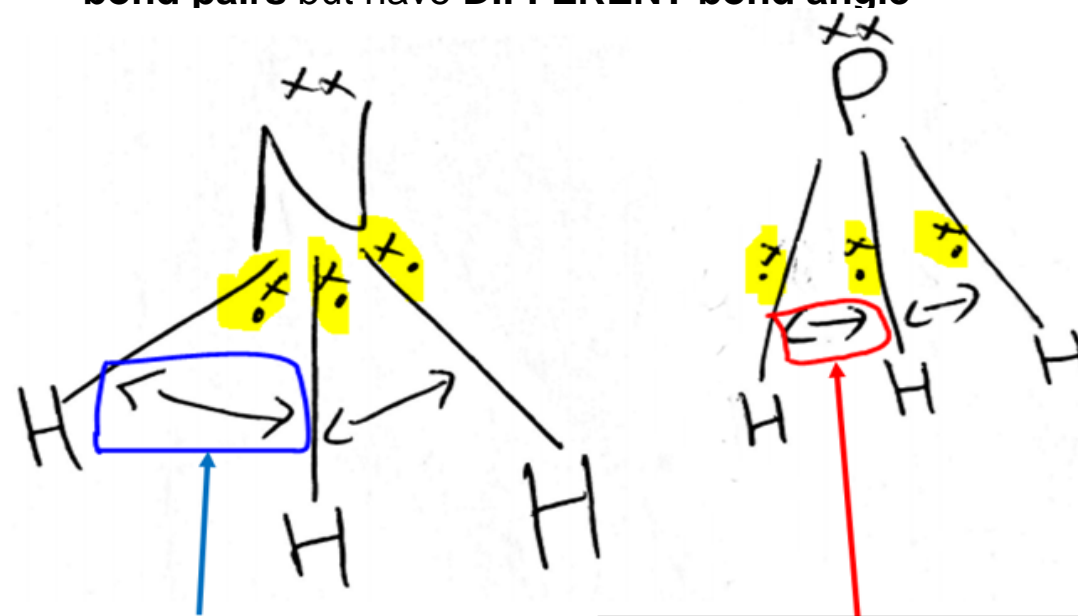
**Principle 3:** Repulsion between bond pairs of electrons increases with an increase in the electronegativity of the central atom

# Shapes and Bond Angles of Molecules

**Principle 1 & 2** is use to explain the shape & bond angle for molecules with **different number of lone pairs & bond pairs**



**Principle 3** is use to explain exception cases where molecules with the **SAME number of lone pairs and bond pairs** but have **DIFFERENT bond angle**



The bonding electrons are **MORE** strongly attracted to the central atom, hence they are **NEARER** to the neighbouring bonding electrons. This result in **GREATER** repulsion between each other

The bonding electrons are **LESS** strongly attracted to the central atom, hence they are **FURTHER** from the neighbouring bonding electrons. This result in **LESSER** repulsion between each other

# Practice Questions

Question: Determine the number of sigma and pi bonds in each molecules

- (i) ethanoic acid
- (ii) hydrogen cyanide

Answer:

- (i) 7  $\sigma$  and 1  $\pi$  bonds
- (ii) 2  $\sigma$  and 2  $\pi$  bonds

Question: Explain whether  $\text{Na}_2\text{O}$  has more covalent characteristic or  $\text{Al}_2\text{O}_3$  has more covalent characteristic.

Answer:  $\text{Al}_2\text{O}_3$  has more covalent characteristic than  $\text{Na}_2\text{O}$ .  $\text{Al}^{3+}$  cation has a greater charge and smaller radius than  $\text{Na}^+$  cation, hence resulting in a greater charge density of the  $\text{Al}^{3+}$  cation. Both  $\text{Al}_2\text{O}_3$  and  $\text{Na}_2\text{O}$  have the same anion. The electron cloud of  $\text{O}^{2-}$  is thus polarised to a larger extent by  $\text{Al}^{3+}$  cation than  $\text{Na}^+$ , resulting in greater covalent characteristic in  $\text{Al}_2\text{O}_3$ .

# Practice Questions

Question: Determine which atom is the donor atom and which is the acceptor atom in the formation of dative bond between  $\text{NH}_4^+$ .

*Answer: N is the donor atom as it has a lone pair of electrons and H is the acceptor atom as it has an empty 1s orbital.*

Question: Draw the polar covalent bond present (if any) in  $\text{H}_2$  or  $\text{HCl}$ . Explain your answer.

*Answer:  $\text{H}_2$  does not contain polar covalent bond as both H atoms have the same tendency to attract the bonding electrons. Therefore, the shared electron-pair is equidistant between both atoms.  $\text{HCl}$  contains a polar covalent bond due to the difference in electronegativities between hydrogen and chlorine. Cl has a stronger tendency to attract electrons than H and hence the shared electron-pair is closer to the Cl atom,  $\text{H} \rightarrow \text{Cl}$ .*

Question: Determine which of the following are polar or non-polar compounds:  $\text{CO}_2$ ,  $\text{CCl}_4$ ,  $\text{CHCl}_3$ ,  $\text{PF}_3$ ,  $\text{H}_2\text{O}$

*Answer:*

*Polar compounds:  $\text{CHCl}_3$ ,  $\text{PF}_3$ ,  $\text{H}_2\text{O}$*

*Non polar compounds:  $\text{CO}_2$ ,  $\text{CCl}_4$*



# Practice Questions

Question: Which metal has the strongest metallic bond strength?

A: Be      B: Mg      C: Ca      D : Sr

Answer: D

Question: Which have a higher melting point, MgO or NaCl. Explain your answer.

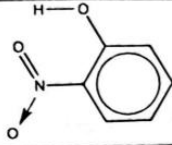
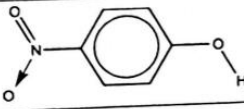
*Answer: Both compounds are ionic compounds with ionic bonding. MgO has a higher melting point than NaCl. Both MgO and NaCl have giant ionic lattice structure. The charge of  $Mg^{2+}$  is greater than that of  $Na^{+}$ , while the charge of  $O^{2-}$  is greater than that of  $Cl^{-}$ . The ionic radius of  $Mg^{2+}$  is smaller than  $Na^{+}$ , while the ionic radius of  $O^{2-}$  is smaller than  $Cl^{-}$ . This indicates that the electrostatic attractions between  $Mg^{2+}$  and  $O^{2-}$  are stronger. Therefore, a larger amount of energy is required to overcome the stronger ionic bonds in MgO, hence resulting in a higher melting point.*

Question: Explain the reason for the difference in boiling point between HF,  $NH_3$  and  $H_2O$ .

*Answer: All three molecules are polar simple covalent structures and are able to form H-Bonds between the molecules.  $H_2O$  has the highest boiling point as it is able to form the most number of H-Bond per molecule. Hence larger amount of energy is required to overcome the higher number of H-Bonds present, resulting in highest boiling point. HF has higher boiling point than  $NH_3$  as the H-F bond has greater magnitude of the dipole moment as compared to N-H bond. Hence larger amount of energy is required to overcome the stronger H-F bond as compared to the N-H bond.*

# Practice Questions

Question: Explain the difference in boiling point between the following compounds:

compound	structure	boiling point / °C	solubility in water / g per 100 g of water
2-nitrophenol		214	0.2
4-nitrophenol		259	1.7

*Answer: 2-nitrophenol and 4-nitrophenol are both polar simple molecular compounds, however 2-nitrophenol have a lower boiling point than 4-nitrophenol. In 2-nitrophenol, due to their close proximity to each other, the  $-OH$  and  $-NO_2$  groups form intramolecular hydrogen bonding. This reduces the number of sites available for intermolecular hydrogen bonding to be formed in 2-nitrophenol.*

*However in 4-nitrophenol, the  $-OH$  and  $-NO_2$  groups are further from each other and hence reduce the probability of forming intramolecular hydrogen bonding. Therefore more intermolecular hydrogen bonding will be formed for the 4-nitrophenol molecules as compared to the 2-nitrophenol molecules. Hence more energy is required to overcome the more extensive intermolecular hydrogen bonding between 4-nitrophenol molecules resulting in 4-nitrophenol molecules having higher boiling point.*

# Practice Questions

Question: Arrange the boiling points of these polar compounds (HCl, HBr, HI) in decreasing order.

Ans: HCl have the highest boiling point, followed by HBr than HI. Electronegativity decrease from Cl to Br to I. Hence greater difference in electronegativity occurs between HCl followed by HBr than HI. Therefore the polarity of the hydrogen halide decrease from HCl to HBr to HI and greatest amount of energy is required to overcome the *pd-pd* interactions between HCl than HBr than HI. This thus explained the decrease in boiling point from HCl to HBr to HI.

Question: Arrange the boiling point of propane, dimethyl ether and ethanol in decreasing order.

Compound	$M_r$	Boiling point ( $^{\circ}\text{C}$ )
$\text{CH}_3\text{CH}_2\text{CH}_3$ , propane	44	-42
$\text{CH}_3\text{OCH}_3$ , dimethyl ether	46	-25
$\text{CH}_3\text{CH}_2\text{OH}$ , ethanol	46	78

Ans: All three molecules are simple covalent structures. Ethanol is able to form strong H-bonds between the molecules, dimethyl ether is able to form *pd-pd* interactions and propane is able to form only *id-id* interactions. Larger amount of energy is required to take in to overcome the strong intermolecular H-Bonding between ethanol molecules as compared to the *pd-pd* interactions between dimethyl ether molecules and the *id-id* interactions between propane molecules. This thus result in ethanol having the higher boiling point as compared to dimethyl ether and propane.

Dimethyl ether has higher boiling point than propane as more energy is required to be taken in to overcome the stronger *pd-pd* interactions present between dimethyl ether molecules as compared to the weaker *id-id* interactions between propane molecules.

# Practice Questions

Question: The table shows the properties of three non-polar simple covalent molecules

Molecule	$M_r$	Boiling point ( $^{\circ}\text{C}$ )
$\text{CH}_3\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_3$ (pentane)	70.0	36.0
$\text{CH}_3\text{CH}(\text{CH}_3)_2\text{CH}_3$ (2,2-dimethylpropane)	70.0	24.5
$\text{CH}_3\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_3$ (hexane)	84.0	68.7

Explain the difference in the boiling point of the three simple covalent molecules.

*Ans: All three molecules are non-polar simple covalent structures and have id-id interactions present between the molecules. Hexane has a larger electron cloud size as compared to the other two molecules, hence its electron cloud is more polarisable. Hence more energy is required to overcome the stronger id-id interactions between hexane molecules, thus hexane has the highest boiling point.*

*Pentane has a higher boiling point than 2,2-dimethylpropane due to the linear structure of pentane as compared to the branched structure of 2,2-dimethylpropane. Therefore pentane has a larger surface area for contact, forming more extensive id-id interactions between its molecules. More energy is thus required to overcome the stronger id-id interactions between pentane molecules as compared to between 2,2-dimethylpropane molecules, resulting in a higher boiling point for pentane as compared to 2,2-dimethylpropane.*



# Practice Questions

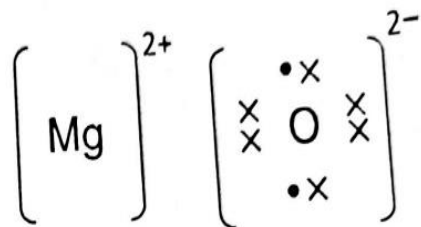
Question: Explain why propanone,  $\text{CH}_3\text{COCH}_3$ , is miscible with water but tetrachloromethane,  $\text{CCl}_4$ , is not.

*Answer:  $\text{CH}_3\text{COCH}_3$  and water have simple molecular structures.  $\text{CH}_3\text{COCH}_3$  molecules can form H-bonding with water molecules (solute-solvent interaction). The energy evolved when H-bonds are formed between  $\text{CH}_3\text{COCH}_3$  molecules and water molecules is sufficient to overcome the energy required to overcome the  $\text{CH}_3\text{COCH}_3$  molecules and the H-bonding between water molecules. Hence  $\text{CH}_3\text{COCH}_3$  molecules are miscible in water.*

*$\text{CCl}_4$  also has a simple molecular structure and is non-polar with id-id interaction between the molecules.  $\text{CCl}_4$  can only forms id-id interactions with water molecules. The energy evolved when id-id interactions are formed between  $\text{CCl}_4$  molecules and water molecules is not sufficient to overcome the energy required to overcome the stronger H-bonding between water molecules. Therefore  $\text{CCl}_4$  molecules are not miscible in water.*

Question: Draw the dot-and-cross diagram of  $\text{MgO}$

Answer:



# Practice Questions

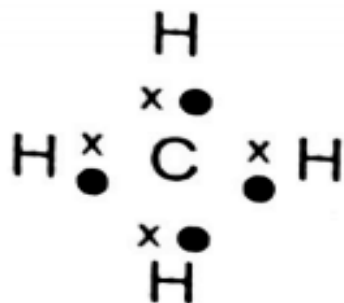
Question: Draw the dot-and-cross diagram of  $\text{CO}_2$ .

Answer:



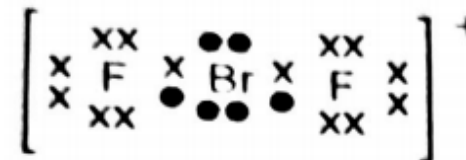
Question: Draw the dot-and-cross diagram of  $\text{CH}_4$ .

Answer:



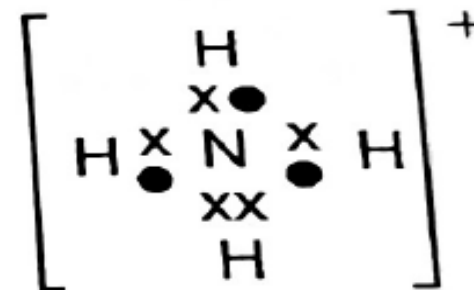
Question: Draw the dot-and-cross diagram of  $\text{BrF}_2^+$

Answer:



Question: Draw the dot-and-cross diagram of  $\text{NH}_4^+$

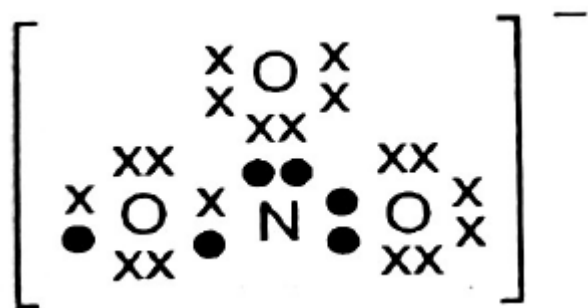
Answer:



# Practice Questions

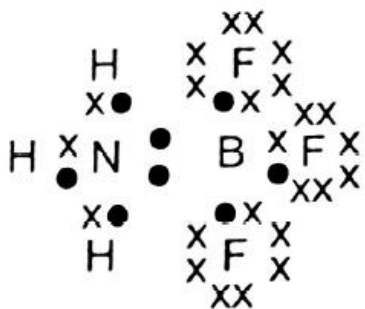
Question: Draw the dot-and-cross diagram of  $\text{NO}_3^-$

Answer:



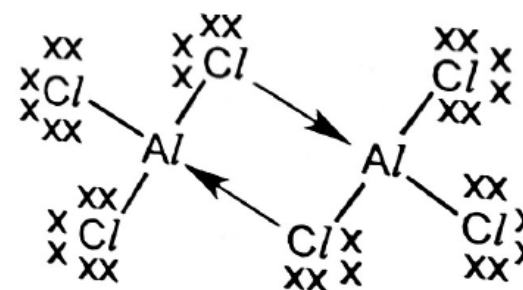
Question: Draw the dot-and-cross diagram for  $\text{BF}_3 \cdot \text{NH}_3$

Answer:



Question: Draw the dot-and-cross diagram for  $\text{Al}_2\text{Cl}_6$

Answer:



Question: Draw the dot-and-cross diagram for  $\text{N}_2$

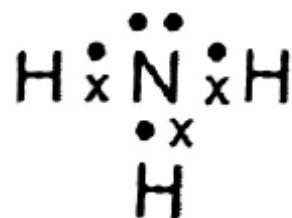
Answer:



# Practice Questions

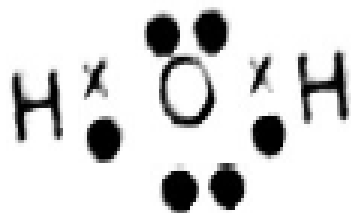
Question: Draw the dot-and-cross diagram for  $\text{NH}_3$

Answer:



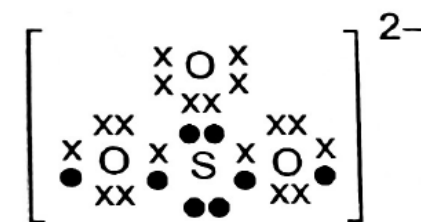
Question: Draw the dot-and-cross diagram for  $\text{H}_2\text{O}$

Answer:



Question: Draw the dot-and-cross diagram for  $\text{SO}_3^{2-}$

Answer:



Question: Draw the dot-and-cross diagram for  $\text{CO}_3^{2-}$

Answer:

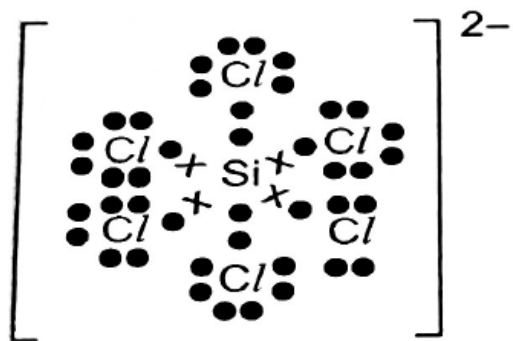




# Practice Questions

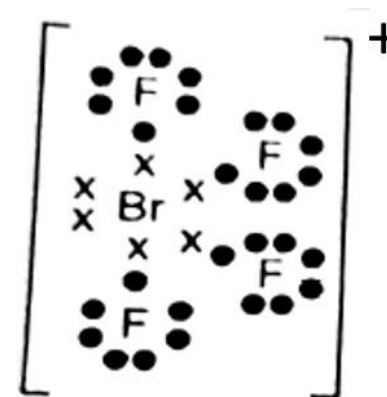
Question: Draw the dot-and-cross diagram for  $\text{SiCl}_6^{2-}$

Answer:



Question: Draw the dot-and-cross diagram for  $\text{BrF}_4^+$

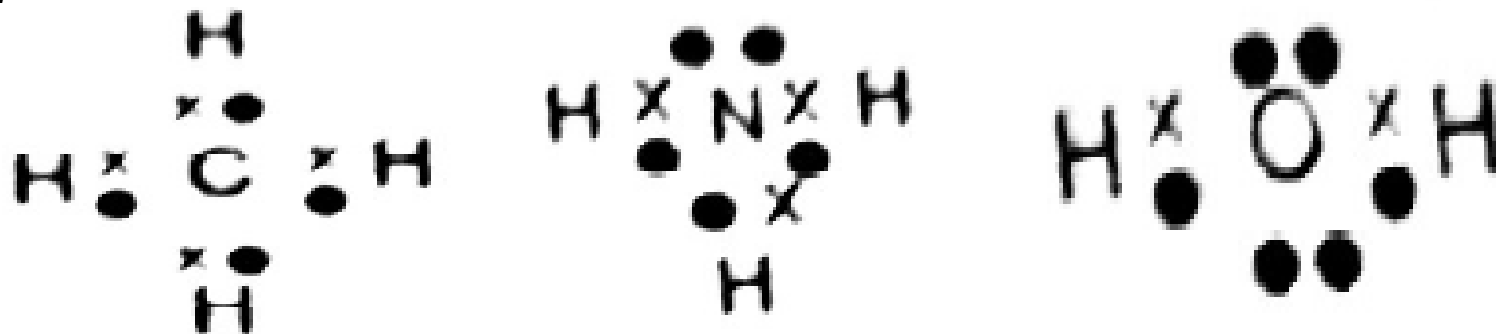
Answer:



# Practice Questions

Question: State and explain the difference in bond angle and shape between  $\text{CH}_4$ ,  $\text{NH}_3$  and  $\text{H}_2\text{O}$ .

Answer:



$\text{CH}_4$ : It is tetrahedral with a bond angle of  $109.5^\circ$ .

$\text{NH}_3$ : It is trigonal pyramidal with a bond angle of  $107^\circ$ .

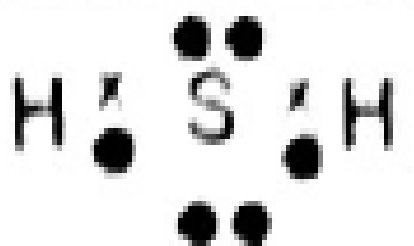
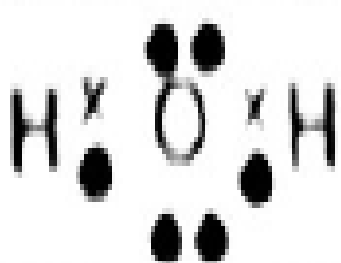
$\text{H}_2\text{O}$ : It is bent with a bond angle of  $104.5^\circ$ .

$\text{H}_2\text{O}$  has the smallest bond angle as it has 2 lone pairs of electrons hence it experience greatest repulsion resulting in it having the smallest bond angle.  $\text{CH}_4$  does not have any lone pairs of electrons thus repulsion only occurs between bond pairs hence it experience the least repulsion and thus the greatest bond angle.  $\text{NH}_3$  only has 1 lone pair of electrons hence it has a bond angle greater than  $\text{H}_2\text{O}$  but smaller than  $\text{CH}_4$ .

# Practice Questions

Question: Explain the difference in bond angle between  $\text{H}_2\text{O}$  and  $\text{H}_2\text{S}$  even though both have the same shape.

Answer:



Both  $\text{H}_2\text{O}$  and  $\text{H}_2\text{S}$  have 2 bond pairs and 2 lone pairs of electrons hence giving them a bent shape with a bond angle below  $109.5^\circ$ . However, since O is more electronegative than S, the bonding electrons in  $\text{H}_2\text{O}$  are drawn nearer to the nucleus than the bonding electrons in  $\text{H}_2\text{S}$ . Greater repulsion thus occurs between the bond pairs of electrons in  $\text{H}_2\text{O}$ , resulting in  $\text{H}_2\text{O}$  having a greater bond angle than  $\text{H}_2\text{S}$ .



# Test yourself!

(a) show understanding that all chemical bonds are electrostatic in nature and describe:

- i. ionic bond as the electrostatic attraction between oppositely charged ions
- ii. covalent bond as the electrostatic attraction between a shared pair of electrons and positively charged nuclei
- iii. metallic bond as the electrostatic attraction between a lattice of positive ions and delocalised electrons

(b) describe, including the use of 'dot-and-cross' diagrams

- i. ionic bonding as in sodium chloride and magnesium oxide
- ii. covalent bonding as in hydrogen; oxygen; nitrogen; chlorine; hydrogen chloride; carbon dioxide; methane; ethene
- iii. co-ordinate (dative covalent) bonding, as in formation of the ammonium ion and in the  $Al_2Cl_6$  molecule

(c) describe covalent bonding in terms of orbital overlap (limited to s and p orbitals only), such as  $\sigma$  and  $\pi$  bond (link with Chapter 11.1: Hydrocarbons)

(d) explain the shapes of, and bond angles in, molecules such as  $BF_3$  (trigonal planar);  $CO_2$  (linear);  $CH_4$  (tetrahedral);  $NH_3$  (trigonal pyramidal);  $H_2O$  (bent);  $SF_6$  (octahedral) by using the Valence Shell Electron Pair Repulsion theory

(e) predict the shapes of, and bond angles in, molecules analogous to those specified in (d)

# Test yourself!

- (f) explain and deduce bond polarity using the concept of electronegativity  
[quantitative treatment of electronegativity is **not** required]
- (g) deduce the polarity of a molecule using bond polarity and its molecular shape (analogous to those specified in (d))
- (h) describe the following forces of attraction (electrostatic in nature):
- intermolecular forces, based on permanent and induced dipoles, as in  $\text{CHCl}_3(\text{l})$ ;  $\text{Br}_2(\text{l})$  and the liquid noble gases
  - hydrogen bonding, using ammonia and water as examples of molecules containing  $-\text{NH}$  and  $-\text{OH}$  groups
- (i) outline the importance of hydrogen bonding to the physical properties of substances, including ice and water
- (j) explain the terms *bond energy* and *bond length* for covalent bonds
- (k) compare the reactivities of covalent bonds in terms of bond energy, bond length and bond polarity

# Test yourself!

(l) describe, in simple terms, the lattice structure of a crystalline solid which is:

- i. ionic, as in sodium chloride and magnesium oxide
- ii. simple molecular, as in iodine
- iii. giant molecular, as in graphite and diamond
- iv. hydrogen-bonded, as in ice
- v. metallic, as in copper

[the concept of the 'unit cell' is **not** required]

(m) describe, interpret and/or predict the effect of different types of structure and bonding on the physical properties of substances

(n) suggest the type of structure and bonding present in a substance from given information



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