KSQUARE CAREER INSTITUTE

CHEMICAL BONDING & MOLECULAR STRUCTURE

Class 11 Chemistry • Kössel–Lewis Approach, Formal Charge, Octet Rule • NEET / JEE Notes

Scan to get all notes

Topics Covered

1. Chemical bond — definition, energetics, types (strong & weak)
2. Theories of bonding — classical vs modern; potential energy curve
3. Kössel–Lewis approach (octet rule), Lewis symbols, ionic & covalent bond formation
4. Writing Lewis dot structures — bond-count formula
5. Formal charge & average formal charge / bond order in oxo-anions
6. Limitations of the octet rule; odd-electron species; Lewis acids & bases
7. Added: ionic bond conditions, lattice enthalpy, ionic vs covalent comparison, practice MCQs

CHEMICAL BOND

Definition

The **force of attraction** between any chemical species like **atoms, ions or molecules** that holds them together.

- **Bond formation** is always **exothermic** (ΔH = –ve) — energy is released.
- **Bond dissociation** is always **endothermic** (ΔH = +ve) — energy is absorbed.
- For bond formation: **attractive force > repulsive force**.

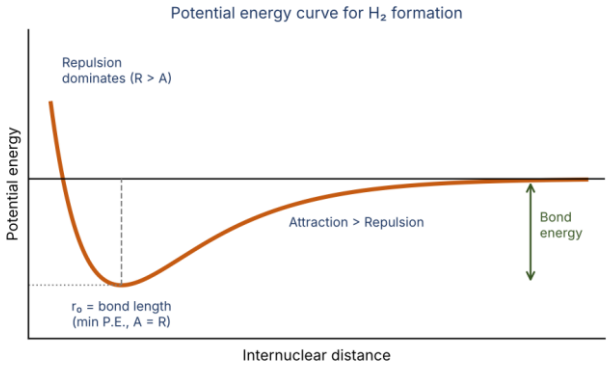
Types of Chemical Bonds

Strong bonds (4 types)	Description	Weak bonds (2 types)
(A) Ionic bond	Between ions, due to **complete transfer of electrons**	**(E) Hydrogen bond**
(B) Covalent bond	Formed by **equal sharing** of electrons	**(F) van der Waals' bond**
(C) Co-ordinate bond	Formed by **unequal sharing** (one atom donates both electrons)	
(D) Metallic bond	Present between **metal–metal** atoms (kernels in a sea of electrons)	

Theories to Explain Chemical Bonding

Concept	Theory	Idea
Classical concept	**Octet rule** (Kössel–Lewis)	Every atom tends to achieve a **stable outer octet** ($8e^-$ in the outermost shell) to gain stability; potential energy ↓
Modern concept	**VBT, VSEPR, MOT**	Every atom tends to form the **maximum possible bonds** to achieve **minimum potential energy** and maximum stability: **stability ∝ 1/P.E.**

Potential Energy Curve



- As two atoms approach, **attraction > repulsion (A > R)** and potential energy decreases.
- At the **equilibrium bond length**, P.E. is **minimum** and **A = R** — this is the **bond-formation point**.
- If they come closer, **repulsion dominates (A < R)** and P.E. rises sharply.
- Energy needed to separate the bonded atoms = **bond dissociation energy**.

KÖSSEL–LEWIS APPROACH (OCTET RULE)

- Based on the **inertness of noble gases** ($ns^2\ np^6$ → $8e^-$ in valence shell = octet complete).
- Atoms combine either by **transfer of valence electrons** (gaining or losing) or by **sharing of valence electrons** to have an **octet**.
- For **H and He**, a **duplet** ($2e^-$ in the valence shell, $1s^2$) is completed.

Lewis Symbols

Show **valence electrons** (electrons in the outermost, n^{th} shell) as dots around the symbol. Valence electrons remain the **same in a group**.

Element	Li	Be	B	C	N	O	F	Ne
Valence config.	$2s^1$	$2s^2$	$2s^22p^1$	$2s^22p^2$	$2s^22p^3$	$2s^22p^4$	$2s^22p^5$	$2s^22p^6$
Valence e^- (dots)	1	2	3	4	5	6	7	8

▸ Octet Completed by Transfer of Electrons — Ionic Bond

Na ($1s^2\ 2s^2\ 2p^6\ 3s^1$) → **Na^+** ($1s^2\ 2s^2\ 2p^6$) + e^-; Cl ($1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$) + e^- → **Cl^-** ($1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$). Both achieve an octet ⇒ **Na^+Cl^- (ionic bond)**.

▸ Octet Completed by Sharing of Electrons — Covalent Bond

Molecule	Lewis structure	Bond pairs (shared e^-)	Lone pairs (non-bonding e^-)	Bond order
Cl_2	:C̈l–C̈l: (Cl–Cl)	1 (2 e^-)	6 (12 e^-)	1
O_2	Ö=Ö (O=O)	2 (4 e^-)	4 (8 e^-)	2
N_2	:N≡N: (N≡N)	3 (6 e^-)	2 (4 e^-)	3

- No. of bond pairs = number of bonds = **bond order** = (no. of shared e^-)/2.
- No. of lone pairs = (no. of unshared e^-)/2.

HOW TO WRITE LEWIS DOT STRUCTURES

Valid only for molecules that **follow the octet rule**.

Number of bonds (bond pairs)

No. of bonds = [Required valence e^- – Available valence e^-] / 2

Required = 8 × (no. of atoms) (use 2 for each H)

Available = sum of valence e^- (+1 for each –ve charge, –1 for each +ve charge)

Species	Required	Available	No. of bonds	Structure
CO	8 × 2 = 16	4 + 6 = 10	(16 – 10)/2 = **3**	C≡O
CN^-	16	4 + 5 + 1 = 10	**3**	$[C≡N]^-$
O_3	8 × 3 = 24	18	(24 – 18)/2 = **3**	O=O–O
CO_2	24	4 + 6 + 6 = 16	**4**	O=C=O
N_3^-	24	5 + 5 + 5 + 1 = 16	**4**	$[N=N=N]^-$

FORMAL CHARGE

Formal charge (F.C.)

The **extra charge** present on an atom in a given Lewis dot structure due to (apparent) gain or loss of electrons.

F.C. = (valence e^- in free atom) – (non-bonding e^-) – ½ (bonding/shared e^-)

Short trick: **F.C. = "jitne hone chahiye – jitne hai"** = valence e^- of free atom – electrons 'owned' by the atom in the structure (lone-pair e^- + one e^- per bond).

Species	Structure	Formal charges
CO_2 (a)	O=C=O	O: 0, C: 0, O: 0 (**preferred**)
CO_2 (b)	$^{+}O{\equiv}C{-}O^{-}$	O: +1, C: 0, O: –1
N_3^- (a)	$^{-}N{=}N^{+}{=}N^{-}$	–1, +1, –1 (**more stable** — symmetrical charge distribution)
N_3^- (b)	$N{\equiv}N^{+}{-}N^{2-}$	0, +1, –2
CO	$^{-}C{\equiv}O^{+}$	C: –1, O: +1
O_3	$O{=}O^{+}{-}O^{-}$	0, +1, –1
CN^-	$^{-}C{\equiv}N$	C: –1, N: 0
SCN^-	$^{-}S{-}C{\equiv}N$	S: –1, C: 0, N: 0
NF_3, CCl_4	Single bonds	All atoms 0
NO_2^-	$O{=}N{-}O^{-}$	O(double): 0, N: 0, O(single): –1
CO_3^{2-}	$O{=}C(-O^{-})_2$	O(double): 0, C: 0, each O(single): –1

NEET Point / PYQ Alert

The most stable Lewis structure has **lowest formal charges** and places **negative charge on the more electronegative atom**.

Formal charge is not the actual charge — it only helps choose the best structure.

Trick — Formal Charge on Terminal Oxygen in Oxo-anions

Oxygen bonded by	Formal charge on O
Single bond (–O)	–1
Double bond (=O)	0
Triple bond (≡O)	+1

Average Formal Charge & Average X–O Bond Order (Resonance)

Formulas

Average F.C. on O = Total charge on oxygen atoms ÷ number of O atoms over which the charge is spread

Average X–O bond order = Total number of X–O bonds ÷ number of X–O pairs participating in resonance

X = central atom; oxo-anion from acid: $H_2CO_3 \rightarrow CO_3^{2-}$, $H_2SO_4 \rightarrow SO_4^{2-}$, $HClO_4 \rightarrow ClO_4^-$

Ion	Formula	Avg F.C. on O	Avg X–O bond order
Carbonate	CO_3^{2-}	–2/3	4/3

Ion	Formula	Avg F.C. on O	Avg X–O bond order
Sulphate	SO_4^{2-}	–2/4 = –1/2	6/4 = 3/2
Perchlorate	ClO_4^-	–1/4	7/4
Sulphite (HW)	SO_3^{2-}	–2/3	4/3
Chlorite (HW)	ClO_2^-	–1/2	3/2
Phosphate (HW)	PO_4^{3-}	–3/4	5/4
Nitrite (HW)	NO_2^-	–1/2	3/2
Chlorate (HW)	ClO_3^-	–1/3	5/3

LIMITATIONS OF THE OCTET RULE

Limitation	Description	Examples
(i) Hypervalent (expanded octet)	Central atom has **more than 8** valence electrons (uses d-orbitals, period 3 onwards)	PCl_5 (10 e^-), SF_6 (12 e^-), H_2SO_4 (12 e^-)
(ii) Hypovalent (incomplete octet)	Central atom has **fewer than 8** electrons	LiCl, $BeCl_2$ / BeI_2 (4 e^-), BF_3, BCl_3, BBr_3, BI_3 (BX_3, 6 e^-)
(iii) Odd-electron species	Total electrons odd — octet cannot be satisfied for all atoms	**NO, NO_2** (N has 7 e^-), ClO_2, ClO_3, CH_3, CF_3
(iv) Noble gas compounds	Noble gases form compounds with O and F	XeF_2, KrF_2, $XeOF_2$, XeO_2F_2, $XeOF_4$, XeF_4, XeF_6
(v) Energy	Silent about the **relative stability / energy** of molecules	—
(vi) Shape	Does **not account for the shape** of molecules	—

Lewis acids & bases

Lewis acid → electron-pair **acceptor** (e.g. BF_3, $AlCl_3$ — electron-deficient).

Lewis base → electron-pair **donor** (e.g. NH_3, H_2O).

Q1. Which is NOT an exception to the octet rule?

(a) BF_3
(b) $SnCl_4$
(c) BeI_2
(d) PCl_5

Answer: (b) — BF_3, BeI_2 — hypovalent; PCl_5 — hypervalent; Sn in $SnCl_4$ has a complete octet.

IONIC BOND — CONDITIONS (ADDITIONAL TOPIC)

- Formation of an ionic bond is favoured by: **low ionisation enthalpy** of the metal, **high (negative) electron gain enthalpy** of the non-metal, and **high lattice enthalpy** of the compound.

- **Lattice enthalpy** = energy required to completely separate one mole of a solid ionic compound into gaseous ions (NaCl = +788 kJ mol^{-1}). It is the main factor that makes ionic compounds stable.

Property	Ionic compounds	Covalent compounds
Bond formed by	Transfer of electrons	Sharing of electrons
Physical state	Crystalline solids	Gases, liquids or soft solids
Melting/boiling point	High	Low
Solubility	Polar solvents (water)	Non-polar solvents
Conductivity	Conduct in molten/aqueous state	Generally non-conductors
Directional nature	Non-directional	Directional (definite shape)

PRACTICE MCQS

Q2. Bond formation is always

(a) Endothermic
(b) Exothermic
(c) Neither
(d) Depends on the bond

Answer: (b)

Q3. The formal charge on the central oxygen atom in O_3 is

(a) 0
(b) +1
(c) –1
(d) +2

Answer: (b) — 6 – 2 – ½(6) = +1.

Q4. The average S–O bond order in SO_4^{2-} is

(a) 1
(b) 1.25
(c) 1.5
(d) 2

Answer: (c) — 6 bonds over 4 S–O pairs.

Q5. Which of the following is an odd-electron molecule?

(a) CO
(b) NO_2
(c) CO_2
(d) N_2

Answer: (b)

Q6. The number of lone pairs in N_2 is

(a) 1
(b) 2
(c) 3
(d) 4

Answer: (b)

Q7. Average formal charge on O in ClO_4^- is

(a) –1
(b) –1/2
(c) –1/4
(d) 0

Answer: (c)

Q8. Which of the following has an expanded octet?

(a) BF_3
(b) SF_6
(c) NH_3
(d) CH_4

Answer: (b) — 12 electrons around S.

QUICK REVISION

- Bond formation exothermic; dissociation endothermic; attraction > repulsion.
- Strong bonds: ionic, covalent, co-ordinate, metallic. Weak: H-bond, van der Waals.
- Classical: octet rule (Kössel–Lewis). Modern: VBT, VSEPR, MOT — stability ∝ 1/P.E.
- No. of bonds = (8 × atoms – available e^-)/2 (2 for H).
- F.C. = V – N – ½B; terminal O: –O (–1), =O (0), ≡O (+1).
- Avg F.C. on O & avg X–O BO: CO_3^{2-} –2/3, 4/3; SO_4^{2-} –1/2, 3/2; ClO_4^- –1/4, 7/4; PO_4^{3-} –3/4, 5/4.
- Octet exceptions: hypervalent (PCl_5, SF_6), hypovalent ($BeCl_2$, BF_3), odd e^- (NO, NO_2), noble gas compounds.

Notes revise 2 times, finish the homework — bonding is the base of chemistry!