Chemistry

Short Questions with Comprehensive Answers

3.1 Why do atoms form chemical bonds?

1. Why do atoms form chemical bonds?
Atoms form bonds to **lower their potential energy** and become **more stable**. Bond formation allows atoms to achieve a more stable electron arrangement (often a complete outer shell), which is energetically favorable.

2. Which discovery helped chemists understand atomic stability?
The discovery of **noble gases** (He, Ne, Ar, Kr, Xe) showed that atoms with **complete outer shells** are unusually stable and mostly unreactive.

3. What is the Duplet Rule?
It states that **atoms with only the first shell** (like H and He) are stable when they have **two electrons** in that shell.

4. What is the Octet Rule?
Atoms (beyond the first shell) tend to be **most stable** when they have **eight electrons** in their outermost shell. Many atoms bond by losing, gaining, or sharing electrons to complete an **octet**.

5. How does the Octet Rule explain sodium's behavior?
Sodium (Na) has one valence electron. It is **energetically easier** for Na to **lose one electron** (forming Na^+) than to gain seven. Losing one electron helps it reach the **nearest noble gas configuration**.

6. How does hydrogen follow the Duplet?
Hydrogen can **gain one electron** to form **hydride (H^-)** or **lose one electron** to form **H^+ (proton)**—both routes move it toward a **stable duplet (2 electrons)** in its first shell.

7. Why are alkali and alkaline earth metals called electropositive?
They **easily lose** their outer electrons to form **cations**, often bonding with **electronegative non-metals** (like O, Cl). This loss helps them attain a stable noble-gas-like configuration.

3.2 Chemical Bond (definition, energy, types)

8. What is a chemical bond?
A **chemical bond** is a **force of attraction** that holds atoms together in a **molecule or compound**.

9. What happens when atoms approach each other?
Two effects compete: **attractions** (electron–nucleus) and **repulsions** (electron–electron and nucleus–nucleus). If attractions **dominate**, the **system's energy decreases**, and a **bond** forms; otherwise, atoms **repel** and separate.

10. What is electronic configuration?
It is the **arrangement of electrons** in **shells and sub-shells** around an atom's nucleus.

11. What main bond types are discussed in this chapter?
Ionic, **Covalent**, and **Coordinate covalent** bonds (with **Metallic bonding** discussed later as a separate section).

3.2.1 Ionic Bond

12. What is an ionic (electrovalent) bond?
An **ionic bond** forms by **complete transfer** of one or more electrons from one atom to another, producing **oppositely charged ions** (cations and anions) that attract each other. This transfer helps atoms reach **noble gas configurations**.

13. Describe the ionic bonding in sodium chloride (NaCl).
Sodium transfers its **one valence electron** to **chlorine**. Sodium becomes **Na^+** and chlorine becomes **Cl^-**. The **electrostatic attraction** between Na^+ and Cl^- forms the **ionic bond** in NaCl.

14. Which electrons participate in chemical reactions and bonding?
Only **valence (outermost shell) electrons** participate.

15. What is meant by a crystal lattice in ionic compounds?
A **crystal lattice** is a **3-D ordered arrangement** of ions where each ion is **surrounded by oppositely charged ions**, maximizing attraction and stability.

16. Give other examples of ionic compounds mentioned.
KCl, **NaF**, **KBr**, **MgF_2**, **CaF_2**, and **$CaCl_2$** (formed when Ca loses two electrons and each Cl gains one).

17. How is calcium chloride ($CaCl_2$) formed ionically?
Calcium **loses two electrons** to form **Ca^{2+}**; two chlorine atoms **each gain one electron** to form **$2Cl^-$**. The lattice is built from **Ca^{2+}** and **Cl^-** ions.

18. What kinds of elements usually form ionic bonds?
Typically a **metal** (especially alkali/alkaline earth) and a **non-metal** (e.g., halogens, oxygen) where **electron transfer** is feasible.

3.2.2 Covalent Bond (formation, single/double/triple, polarity)

19. What is a covalent bond?
A **covalent bond** forms when two atoms **mutually share** one or more **pairs of electrons**. One shared pair makes a **single bond**.

20. Why does sharing electrons lower energy in covalent bonding?
As atoms approach, the **attraction** between each nucleus and the other atom's electrons **reduces energy**; at an optimal distance where attractions exceed repulsions, a **stable molecule** with **minimum energy** forms.

21. How are single, double, and triple covalent bonds represented?
A **single bond**: **–** (one shared pair), **double bond**: **=** (two shared pairs), **triple bond**: **≡** (three shared pairs).

22. Give examples showing single, double, and triple bonds.

- **H_2O (water)**: two **single** O–H bonds (each H shares one electron with O).
- **CO_2**: two **double** C=O bonds (each O shares two electrons with C).
- **HCN**: H–C≡N has a **triple** C≡N bond.

23. Describe covalent bonding in ammonia and methane.

- **NH_3 (ammonia)**: N shares **three** pairs (one with each H) → three N–H **single bonds**.
- **CH_4 (methane)**: C shares **four** pairs (one with each H) → four C–H **single bonds**.

24. What is a non-polar covalent bond? Give examples.
A **non-polar** covalent bond forms when **identical atoms** share electrons **equally**; the electron pair lies midway. Examples: **H_2**, **Cl_2**.

25. What is a polar covalent bond? Give examples.
A **polar** covalent bond forms between **different atoms** where the shared electrons are **pulled toward the more electronegative atom**, creating **partial charges (δ^+/δ^-)**. Examples: **HCl** (δ^+H—Clδ^-), **H_2O** (polar O–H bonds).

3.2.3 Coordinate Covalent Bond (dative bond)

26. What is a coordinate covalent (dative) bond?
It is a **covalent bond** where the **shared pair** is **donated by one atom** (the **donor**) to an **acceptor** with an empty orbital. It is shown by an **arrow (→)** from donor to acceptor.

27. How does the hydronium ion (H_3O^+) form a coordinate bond?
In water, oxygen has **lone pairs**. A **proton (H^+)** accepts a **lone pair from O**, forming **H_3O^+** with a coordinate bond **O→H**. After formation, all three O–H bonds behave **identically**.

28. Explain the reaction between ammonia (NH_3) and boron trifluoride (BF_3).
NH_3 (donor) uses its **lone pair on N** to fill **B's vacant orbital** in **BF_3 (acceptor)**, forming a **coordinate bond (N→B)** and a stable adduct.

29. Give more examples of coordinate bonding from the chapter.
Ammonium ion (NH_4^+) forms when **NH_3 donates** a pair to **H^+**; **protonated ethyl alcohol** forms when **ethanol's oxygen** donates a pair to **H^+**.

30. After a coordinate bond forms, does it differ from an ordinary covalent bond?
No. Once formed, there is **no functional difference**; the bonds are **indistinguishable**. On **bond breaking**, however, **both electrons return to the original donor**.

3.3 Metallic Bond (properties, model, comparisons)

31. What is a metallic bond?
A metallic bond is the attraction between **positively charged metal ions** arranged in a lattice and a **"sea" of mobile (delocalized) valence electrons** that move freely throughout the metal.

32. Why do metals conduct heat and electricity well?
Because their **delocalized electrons** can **move freely**, carrying **thermal** and **electrical energy** efficiently through the lattice.

33. Why are metals malleable and ductile?
Metal atoms are arranged in **layers (rows)**. When pressure is applied, layers **slide** over each other **without breaking** the metallic bonding, allowing metals to be **hammered into sheets (malleable)** and **drawn into wires (ductile)**.

34. What factors affect the strength of metallic bonds?
(1) The **charge on metal ions** and (2) the **number of delocalized electrons** per atom. More charge/electrons → **stronger** metallic bonding.

35. Compare metallic bonding in sodium and magnesium.
Na provides **one** electron per atom; **Mg** provides **two** and forms **Mg^{2+}**. Thus **Mg** has **stronger metallic bonding** and a **higher melting point** than Na.

36. Briefly compare metallic and ionic bonds.
Metallic: cations in a lattice bound by **delocalized electrons**; good conductors in **solid** state.
Ionic: **cations–anions** held by **electrostatic forces**; conduct **when molten or in solution**, not as solids.

3.4 Electropositive Character of Metals

37. What is the electropositive character of metals?
It is the **tendency of metals to lose electrons** and form **cations**. Higher ease of losing electrons → **more reactive** metal.

38. How do alkali and alkaline earth metals illustrate electropositivity?
Alkali metals (Na, K) lose electrons **very easily**, reacting vigorously with **water, halogens**, and **acids**. **Alkaline earth metals (Mg, Ca)** lose electrons **less easily**, so their reactions are **less vigorous**.

39. How does aluminum behave with respect to electropositivity?
Aluminum is also **highly electropositive**; it **readily reacts with mineral acids** to produce **salts and hydrogen gas**.

3.5 Electronegative Character of Non-metals

40. What is meant by electronegative character of non-metals?
Non-metals have an **affinity for electrons**; they tend to **gain electrons** and form **anions**, so they are called **electronegative** elements.

41. Which elements are highly electronegative as mentioned?
Fluorine (highest), followed by **oxygen**, **nitrogen**, and **chlorine**.

42. How do non-metals bond with metals and with other non-metals?
With **metals**: they form **ionic compounds** (electron transfer).
With **non-metals**: they form **molecular (covalent) substances** (electron sharing).

3.6 Comparing Properties of Ionic and Covalent Compounds

43. What structural feature characterizes ionic compounds?
They consist of **oppositely charged ions** arranged in a **crystalline lattice**; the overall compound is **electrically neutral** with strong **electrostatic attractions**.

44. Why do ionic compounds have high melting/boiling points? Give a data point.
Strong **ionic attractions** require **much energy** to break. Example: **NaCl melts at 801 °C**.

45. How do ionic compounds behave in solvents and electricity?
They are generally **soluble in polar solvents** (like water). They **conduct electricity** in the **molten state** or in **aqueous solution** due to **mobile ions**.

46. How do covalent compounds exist and what are their typical physical properties?
They mostly exist as **neutral molecules**. Lower molecular mass ones are **gases or low-boiling**

liquids; higher ones can be **solids**. Generally, they have **lower melting/boiling points** than ionic compounds.

47. What is the solubility and electrical behavior of covalent compounds?
They are usually **insoluble in water** but **soluble in non-polar solvents** (e.g., ether/benzene/acetone). They are **poor conductors** of electricity.

3.7 Intermolecular Forces of Attraction

48. What are intermolecular forces? How do they compare to bonds?
They are **attractive forces between molecules**, much **weaker** than the **intramolecular** (bonding) forces that hold atoms together within a molecule.

49. How do intermolecular forces affect melting/boiling points?
Stronger intermolecular attractions → **higher** melting and boiling points; weaker attractions → **lower** melting/boiling points.

50. What are dipole–dipole forces? Give an example.
Attractions between **polar molecules**, where the **δ^+** end of one attracts the **δ^-** end of another. Example: **HCl–HCl** interactions increase its **boiling point** relative to non-polar molecules.

51. What is hydrogen bonding and when does it occur?
A strong type of dipole–dipole attraction occurring when **H** is covalently bonded to **highly electronegative atoms** (**F, O, N**). The hydrogen of one molecule is attracted to a lone pair on the electronegative atom of another molecule.

52. Why does water have an unusually high boiling point?
Because **H_2O molecules** form **extensive hydrogen bonding**, requiring extra energy to separate molecules compared with similar hydrides like **H_2S** and **NH_3**.

3.8 Nature of Bonding and Properties (Ionic vs Covalent Solids; Solutions; Reactions)

53. Describe the bonding and structure in ionic solids.
Ions are held in a **rigid crystal lattice** by **strong, non-directional** electrostatic forces. This makes ionic solids **stable** and generally **high-melting**.

54. Why are ionic solids brittle?
When a **shearing force** shifts layers, **like-charged ions** can be forced closer, causing **strong repulsion** and **cleavage**—so ionic crystals **break easily**.

55. Do ionic solids conduct electricity as solids? Explain.
No. In the solid state, ions are **fixed** in place and cannot move. They conduct **when molten** or **dissolved** because **ions become mobile**.

56. What is hydration of ions and how does it aid solubility?
In water, **polar H_2O molecules** surround ions (**hydration**), stabilizing them in solution and **weakening ionic attractions**, thus **promoting solubility**.

57. Give an example of an ionic reaction in aqueous solution from the chapter.
Mixing **NaCl(aq)** and **$AgNO_3$(aq)** forms **AgCl(s)** as a **white precipitate**, while **Na^+(aq)** and **NO_3^-(aq)** remain in solution.

58. How is electrolysis used with ionic compounds in the chapter examples?

- **Molten NaCl** electrolysis → **sodium metal** and **chlorine gas**.
- **Aqueous NaCl** electrolysis → **sodium hydroxide (NaOH)** and **chlorine gas**.

59. How do covalently bonded elements typically exist on the right side of the periodic table?
As **diatomic molecules** (excluding noble gases): **N_2, O_2, F_2, Cl_2**—with **low densities** and **low boiling points** due to weak intermolecular forces.

60. Which covalent element is liquid at room temperature and which exist as covalent solids?
Bromine (Br_2) is a **volatile liquid**. **Carbon, phosphorus, sulfur** can exist as **covalent solids**.

61. Contrast diamond and graphite (both carbon).

- **Diamond**: each C forms **four covalent bonds** in a **3-D network** → **extreme hardness**, **high melting point**, **electrical insulator**; used for **cutting/polishing/drilling**.
- **Graphite**: **layered** hexagonal sheets with **mobile electrons** between layers → **soft/lubricant**, **conducts electricity**; used in **pencils**, **polishes**, **electrodes**.

62. List common binary covalent gases at room temperature from the chapter.
CH_4, NH_3, H_2S, HCl, NO_2, CO_2, SO_2 are **gases** at room temperature.

63. Which small covalent molecules are liquids at room temperature here, and why?
Water (H_2O) and **hydrogen fluoride (HF)** are **liquids** because of **strong hydrogen bonding** that raises their boiling points.

64. How do some covalent acids behave in water according to the chapter?
Molecules like **HCl**, **H_2SO_4**, and **HNO_3 ionize completely in water**, acting as **very strong acids** and conducting electricity in aqueous solution.