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IB DIPLOMA CHEMISTRY · FIRST ASSESSMENT 2025

Structure 2.2
The Covalent Model

Booklet 1 of 2 · Understandings 2.2.1 – 2.2.3

Covalent bonds, Lewis formulas, multiple bonds & coordination bonds

Complete revision notes & memory points · Standard Level & Higher Level

How the level tags work in this booklet

SL + HL Core material. Every student must know this.

HL Extension material. Understandings 2.2.1–2.2.3 are common to SL and HL, but HL papers reach into this material through the AHL parts of Structure 2.2 itself (expanded octets, sigma and pi bonds) and through Reactivity 3.4 (Lewis acids and bases, transition element complexes). See section 1.

Inside this booklet

1. What the syllabus actually asks of you
2. 2.2.1 — Covalent bonds, the octet rule & Lewis formulas
3. 2.2.2 — Single, double & triple bonds
4. 2.2.3 — Coordination bonds
5. Bringing the three together
6. Key points to memorise
7. Check yourself
8. Answers

1 What the syllabus actually asks of you

Structure 2.2 is built around a single guiding question: *what determines the covalent nature and properties of a substance?* It is the largest subtopic in the Structure strand, so it has been split across two booklets. This first booklet covers the three understandings that describe the covalent bond itself. Everything in Booklet 2 — molecular shape, polarity, intermolecular forces and the HL bonding theories — is built on the Lewis formulas you learn to draw here.

Understanding	What it says	What you must be able to do	Level
2.2.1 Covalent bonds & Lewis formulas	A covalent bond is the electrostatic attraction between a shared pair of electrons and the positively charged nuclei. The octet rule describes the tendency of atoms to gain a valence shell of eight electrons.	Deduce Lewis formulas of molecules and ions with up to four electron pairs on each atom. Show every valence electron, bonding and non-bonding. Cover species with fewer than an octet, using both organic and inorganic examples.	SL+HL
2.2.2 Multiple bonds	Single, double and triple bonds involve one, two and three shared pairs of electrons respectively.	Explain the relationship between the number of bonds, bond length and bond strength.	SL+HL
2.2.3 Coordination bonds	A coordination bond is a covalent bond in which both electrons of the shared pair come from the same atom.	Identify coordination bonds in compounds. At HL this includes transition element complexes.	SL+HL

The SL / HL question, answered properly

All three understandings in this booklet are examined at both levels. HL examiners can still build on them, though, through four routes. Each one is flagged with an HL panel where it arises:

- **Transition element complexes** such as $[Fe(H_2O)_6]^{3+}$ — the HL guidance for 2.2.3 and → *AHL Reactivity 3.4* (Lewis acids and bases).
- **Expanded octets** — central atoms with five or six electron domains, such as PCl_5 and SF_6 → *AHL Structure 2.2.13*. At SL you will never be asked for more than four electron pairs on any atom.
- **Sigma and pi bonds** — why a double bond is not twice as strong as a single bond → *AHL Structure 2.2.15*.
- **Formal charge** — choosing between several possible Lewis formulas for the same species → *AHL Structure 2.2.14* (covered in Booklet 2).

Linking questions the guide attaches to 2.2.1 – 2.2.3

- **Nature of Science:** what are some of the limitations of the octet rule? → section 2.9
- **Structure 1.3:** why do noble gases form covalent bonds less readily than other elements? → section 2.5
- **Structure 2.1:** why do ionic bonds only form between different elements, while covalent bonds can form between atoms of the same element? → section 2.1
- **Reactivity 2.2:** how does the presence of double and triple bonds influence the reactivity of a molecule? → section 3.5
- HL **Reactivity 3.4:** why do Lewis acid–base reactions lead to the formation of coordination bonds? → section 4.5

What the data booklet gives you — and what it does not

- **Given:** the periodic table, from which you read the number of valence electrons; covalent bond lengths (**Section 11**); bond enthalpies and average bond enthalpies (**Section 12**); electronegativity values (Section 9).
- **Not given:** any Lewis formula, the octet rule exceptions, or which atom donates in a coordination bond. These come from understanding, not from lookup.

2 Structure 2.2.1 — The covalent bond, the octet rule and Lewis formulas

SL + HL

2.1 Why non-metals share rather than transfer

In Structure 2.1, ions formed because one atom (a metal, low ionization energy) could give up electrons and the other (a non-metal, high effective nuclear charge) could take them. When two **non-metals** meet, neither condition is met. Both atoms hold their valence electrons strongly and have similar electronegativities, so neither can pull an electron completely away from the other. Instead, each atom contributes an electron to a pair that both of them share. Each atom can then count the shared pair towards its own valence shell, and both move closer to a noble gas configuration.

This answers the linking question from Structure 2.1. An ionic bond needs a **large difference** in electronegativity, which can only exist between two different elements. Two atoms of the same element have **identical** electronegativities, so the only way they can bond is by sharing. That is why H_2, Cl_2, O_2 and N_2 are all covalent.

Definitions to learn word-for-word

- **Covalent bond** — the electrostatic attraction between a shared pair of electrons and the positively charged nuclei of the bonded atoms.
- **Molecule** — a discrete group of atoms held together by covalent bonds.
- **Octet rule** — the tendency of atoms to gain a valence shell containing a total of eight electrons.
- **Bonding pair** — a pair of electrons shared between two atoms.
- **Non-bonding pair (lone pair)** — a pair of valence electrons that belongs to one atom and is not involved in bonding.
- **Lewis formula** — a diagram that shows *all* the valence electrons in a covalently bonded species, both bonding and non-bonding pairs.

The words *valence electrons* and *outer-shell electrons* mean the same thing and can be used interchangeably.

2.2 Where the attraction actually is

The definition says the bond is an attraction between the shared pair and *the nuclei* — not between the two atoms, and not between the electrons. The shared pair spends most of its time in the region **between** the two nuclei. There it is attracted by both positive nuclei at once, and this holds the nuclei together.

The definition does *not* say "the sharing of electrons". Sharing describes how the bond forms. The bond itself is an electrostatic attraction, and examiners award the mark for that word.

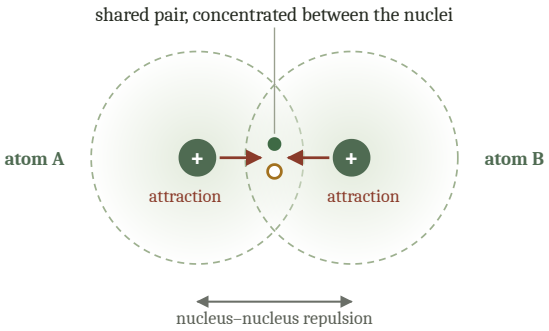


Figure 1 — Both nuclei attract the shared pair that lies between them.

2.3 The energy picture: why a bond has a fixed length

Imagine two hydrogen atoms approaching each other from far apart. The graph tracks the potential energy of the pair as the distance between their nuclei shrinks.

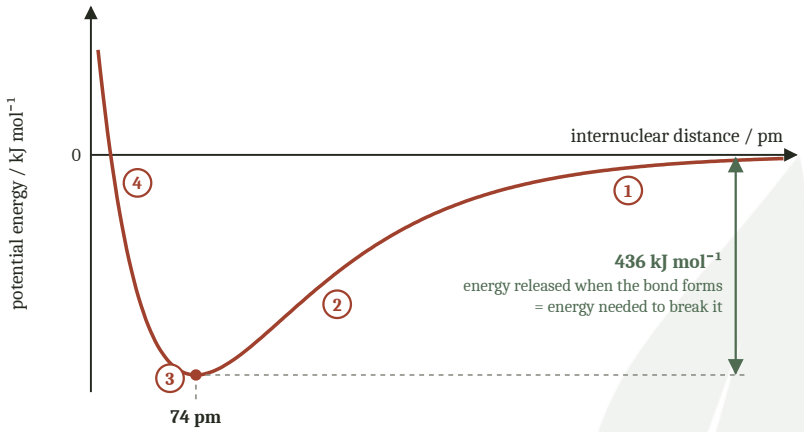


Figure 2 — Energy of two hydrogen atoms against internuclear distance. Values from data booklet Sections 11 and 12.

1. **Far apart.** The atoms do not interact. Their energy is taken as zero.
2. **Approaching.** Each electron starts to feel the pull of the *other* nucleus. The electron density shifts into the region between the nuclei and the energy falls.
3. **Minimum energy.** Attraction and repulsion balance exactly. This separation, 74 pm, is the **bond length**, and the depth of the well, 436 kJ mol^{-1}, is the **bond enthalpy**.
4. **Too close.** Repulsion between the two nuclei (and between the electrons) takes over and the energy climbs steeply.

One sentence to carry into the exam

A covalent bond forms at the separation where the attractions between the shared electrons and both nuclei are balanced by the repulsions between the nuclei, which is the point of minimum energy. Energy is **released** when a bond forms and must be **supplied** to break it.

2.4 The octet rule — and the duet for hydrogen

Sharing gives each atom a share of extra electrons. Atoms in periods 2 and 3 usually keep forming bonds until they have eight electrons in their valence shell, the configuration of the next noble gas. In the chlorine molecule, for example, each chlorine contributes one electron to the bond and keeps three non-bonding pairs. Counting the shared pair, each atom has eight electrons, the configuration of argon.

Hydrogen is the important exception from the start. Its valence shell is the first energy level, which holds only two electrons, so hydrogen is complete with one shared pair: a **duet**, matching helium. Hydrogen therefore always forms exactly one covalent bond and never carries a lone pair.

The octet rule lets you predict how many bonds each atom normally forms. That is simply the number of electrons it is short of a full shell:

Group	1 (H)	14	15	16	17
Valence electrons	1	4	5	6	7
Electrons needed for a full shell	1	4	3	2	1
Usual number of covalent bonds	**1**	**4**	**3**	**2**	**1**
Lone pairs when it forms that many bonds	0	0	1	2	3
Examples	H in all compounds	C in CH_4, Si in $SiCl_4$	N in NH_3, P in PCl_3	O in H_2O, S in H_2S	F, Cl, Br, I

Memorise the bottom two rows. They let you sketch the skeleton of almost any molecule at sight, and in 2.2.3 they become the tool for *spotting* coordination bonds: an atom with more bonds than its usual number has usually received a donated pair.

2.5 Why the noble gases rarely bond

The atoms of every element in group 18 already have a complete valence shell — eight electrons, or two for helium — without sharing anything. Their first ionization energies are the highest in their periods, so their electrons are held very tightly. With nothing to gain from sharing, they show very low reactivity and rarely form covalent bonds. This answers the Structure 1.3 linking question.

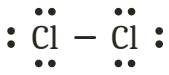
"Rarely" is not "never". Xenon is a large atom with a relatively low ionization energy, and it forms compounds such as XeF_2 and XeF_4 with the most electronegative elements. Xenon has more than eight electrons around it in these compounds, so they are expanded-octet species HL (2.2.13).

2.6 Lewis formulas: the notation

A Lewis formula shows every valence electron in the species. Electron pairs can be drawn as **dots**, **crosses**, a **combination of dots and crosses**, or a **line** for each pair. In this booklet, dots and crosses show which atom each electron originally came from. That is useful for learning, but every electron is identical, so it is not compulsory. The line notation is quicker in exams. Use whichever you prefer, but recognise all of them, and be clear and consistent: a scatter of loose dots will not earn the mark.



dots and crosses



lines and dots



water — dots and crosses



water — lines and dots

The most common lost mark in this subtopic

A structure that shows only the bonds is a **structural formula**, not a Lewis formula. If the question says "Lewis formula", you must draw **every lone pair** on every atom, including all three pairs on each halogen. Missing lone pairs is the single most frequent reason a Lewis formula scores zero.

2.7 Drawing a Lewis formula: the method

1. **Count the valence electrons.** Multiply each element's valence electrons (read from its group) by the number of atoms of that element, and add the results. For an anion, add one electron for each negative charge; for a cation, subtract one for each positive charge.
2. **Build the skeleton.** Place the central atom, which is usually the one that forms the most bonds and is never hydrogen. Join each outer atom to it with one bonding pair.
3. **Complete the octets.** Add the remaining electrons as lone pairs, starting with the *outer* atoms, until each has eight (hydrogen two). Place any electrons left over on the central atom.
4. **Check.** Does the total number of electrons drawn equal your count from step 1? If the central atom is still short of an octet, convert lone pairs on a neighbouring atom into extra shared pairs to make multiple bonds (section 3). For an ion, put the whole structure in square brackets with the charge outside, at the top right.

Worked example — hydrogen peroxide, H_2O_2

1. Count: (2 × 1) + (2 × 6) = **14** valence electrons, which is 7 pairs.

2. Skeleton: oxygen forms two bonds and hydrogen one, so the only arrangement that works is H–O–O–H. That uses 3 bonding pairs (6 electrons).

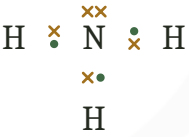
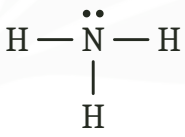
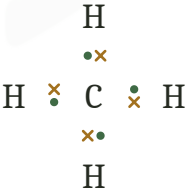
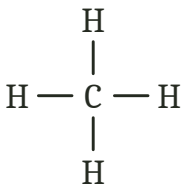
3. Complete: 14 − 6 = 8 electrons remain, which is 4 lone pairs. Each oxygen needs two lone pairs to reach eight (two bonding pairs + two lone pairs). Each hydrogen is already complete.

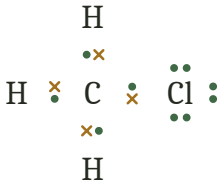
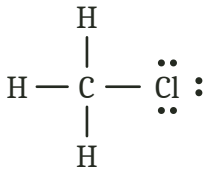
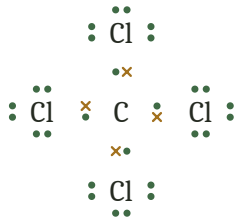
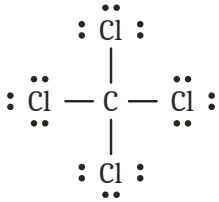
4. Check: 3 bonding pairs + 4 lone pairs = 7 pairs = 14 electrons ✓





2.8 A gallery of molecules you should be able to draw instantly

Species	Valence electrons	Dots and crosses	Lines	Notes
H_2	2 × 1 = 2			Each H has a duet.
HCl	1 + 7 = 8			Cl keeps three lone pairs.
NH_3	5 + (3 × 1) = 8			One lone pair on N — the lone pair that makes NH_3 a base.
CH_4	4 + (4 × 1) = 8			No lone pairs on C.

Species	Valence electrons	Dots and crosses	Lines	Notes
CH_3Cl	4 + 3 + 7 = 14			An organic example: carbon forms four bonds.
CCl_4	4 + (4 × 7) = 32			Twelve lone pairs — all of them must be drawn.

2.9 Lewis formulas of ions

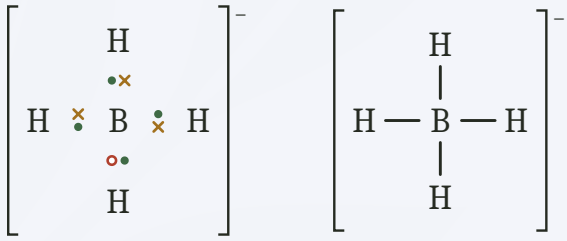
The method is the same, with two additions. Adjust the electron count for the charge in step 1. Then enclose the structure in square brackets with the charge written outside. In the dot-and-cross versions below, the electron gained because of the negative charge is drawn as an open red circle so you can see where it went.

Worked example — the hydroxide ion, OH^-, and the tetrahydridoborate ion, BH_4^-

OH^-: 6 + 1 + 1 (charge) = **8** electrons. One O–H bonding pair leaves three lone pairs on oxygen, giving it a full octet.



BH_4^-: 3 + (4 × 1) + 1 (charge) = **8** electrons, all used in four B–H bonds. The extra electron gives boron its fourth bond and a full octet.



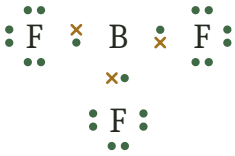
The same method gives a valid Lewis formula for larger ions too. Sulfate, SO_4^{2-}, has 6 + (4 × 6) + 2 = 32 electrons and can be drawn with four S–O single bonds and three lone pairs on each oxygen. HL At HL you will meet a competing structure with an expanded octet on sulfur, and formal charge (2.2.14) is the tool used to decide between them.

2.10 When the octet rule fails: incomplete octets

A small number of atoms form stable molecules in which the central atom has **fewer than eight** valence electrons. This is called an **incomplete octet**. The usual examples are **hydrogen** (always two), **beryllium**, **boron** and **aluminium**. These atoms have only two or three valence electrons to share, and their small size limits how many atoms can crowd round them, so they cannot reach eight by sharing alone.



$BeCl_2$: 2 + 14 = 16 electrons · Be has only 4



BF_3: 3 + 21 = 24 electrons · B has only 6

Species with an incomplete octet have a vacancy in the valence shell of the central atom. That vacancy explains their chemistry: BF_3 and $AlCl_3$ readily accept an electron pair from another species. This is the starting point of coordination bonding, so keep it in mind for section 4.

Two other kinds of species break the rule. **Radicals** have an odd number of electrons, so at least one atom must have fewer than eight; nitrogen monoxide, NO, with 11 valence electrons, is the standard example. HL **Expanded octets** occur when atoms of period 3 and beyond have *more* than eight electrons around them, as in PCl_5 (10) and SF_6 (12). The usual explanation is that these atoms have an empty d sublevel in the same shell that can hold extra electron domains. Period 2 elements have no 2d sublevel, so carbon, nitrogen, oxygen and fluorine can never exceed eight. Expanded octets are AHL 2.2.13 and are covered in Booklet 2.

Nature of Science — what are the limitations of the octet rule?

The octet rule is a **generalisation**, not a law. It describes the behaviour of most compounds of carbon, nitrogen, oxygen and the halogens very well, but it does not explain *why* eight is special, and it breaks down for hydrogen, beryllium, boron and aluminium (too few electrons), for radicals (an odd number of electrons), and for many period 3 and heavier elements (expanded octets). It says nothing at all about transition elements. Scientific rules are judged by how useful they are, and part of using one well is knowing where it stops working. That is why the exceptions are examined as often as the rule itself.

Examiner traps in 2.2.1

- Omitting lone pairs on the outer atoms, especially the three pairs on each halogen.
- Defining a covalent bond as "atoms sharing electrons" with no mention of *electrostatic attraction* and *nuclei*.
- Putting hydrogen in the centre of a skeleton, or drawing it with two bonds.
- Forgetting to adjust the electron count for an ion's charge, or leaving off the square brackets and charge.
- "Completing" the octet on Be in $BeCl_2$ or B in BF_3 by inventing double bonds. Draw them with incomplete octets.

3 Structure 2.2.2 — Single, double and triple covalent bonds SL + HL

3.1 When one shared pair is not enough

Sometimes, after step 3 of the method, an atom is still short of an octet and there are no electrons left to give it. The fix is to share **more than one pair** between the same two atoms. A neighbouring atom turns one of its lone pairs into a second, or third, bonding pair.

Definitions

- **Single bond** — one shared pair of electrons (2 electrons).
- **Double bond** — two shared pairs of electrons (4 electrons).
- **Triple bond** — three shared pairs of electrons (6 electrons).

In line notation, one line represents one shared pair, so = is a double bond and ≡ is a triple bond.

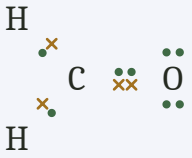
Worked example — methanal, HCHO

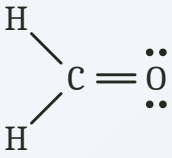
1. Count: 4 + (2 × 1) + 6 = **12** valence electrons (6 pairs).

2. Skeleton: carbon is central, bonded to two H atoms and to O. That uses 3 pairs, leaving 3 pairs.

3. Complete: all 3 remaining pairs go on the outer oxygen, which now has 8. Carbon, however, has only 6: three bonding pairs.

4. Fix: move one oxygen lone pair into the C–O region. It becomes a second shared pair, so C=O is a double bond. Now C has 8, O has 8 (two bonding pairs + two lone pairs), and the total is still 12 ✓





A shortcut worth knowing — the bond-counting check

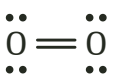
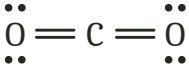
For any species in which **every** atom obeys the octet rule (duet for H):

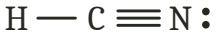
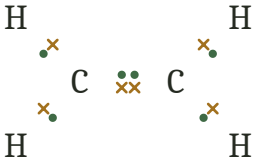
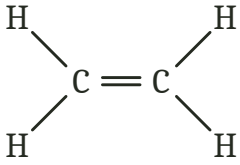
shared electrons = (electrons needed for every atom to be full) − (valence electrons available)

number of bonds = shared electrons ÷ 2

For CO_2: needed = 3 × 8 = 24; available = 4 + 12 = 16; shared = 8, so there are **4 bonds** in total. With two C–O links, that must be two double bonds, O=C=O. For HCN: needed = 2 + 8 + 8 = 18; available = 10; shared = 8, so there are 4 bonds: H–C plus C≡N. The shortcut does *not* work for incomplete or expanded octets. Try it on BF_3 and it wrongly predicts a B=F double bond.

3.2 Gallery: molecules with multiple bonds

Species	Valence electrons	Dots and crosses	Lines	Notes
O_2	2 × 6 = 12			Double bond; two lone pairs on each O.
N_2	2 × 5 = 10			Triple bond; one lone pair on each N.
CO_2	4 + (2 × 6) = 16			Two double bonds; no lone pairs on C.

Species	Valence electrons	Dots and crosses	Lines	Notes
HCN	1 + 4 + 5 = 10			C≡N triple bond; lone pair on N.
C_2H_4	(2 × 4) + 4 = 12			Ethene: C=C double bond.
C_2H_2	(2 × 4) + 2 = 10			Ethyne: C≡C triple bond.

3.3 Bond length and bond strength

Every covalent bond is characterised by two measurable quantities, both of which are tabulated in the data booklet:

- **Bond length** — the distance between the nuclei of the two bonded atoms (Section 11, in pm).
- **Bond strength** — usually measured as the **bond enthalpy**: the energy needed to break one mole of the bond in gaseous molecules (Section 12, in kJ mol^{-1}). For bonds that occur in many compounds, such as C–H, the booklet gives an *average* value. Bond enthalpies are used in calculations in Reactivity 1.2.

The key idea links the two. A double bond places **four** electrons between the nuclei instead of two, and a triple bond places **six**. More negative charge between the nuclei means a stronger electrostatic attraction for both nuclei. This pulls the nuclei closer together and needs more energy to overcome. So as the number of shared pairs increases, **bond length decreases** and **bond strength increases**.

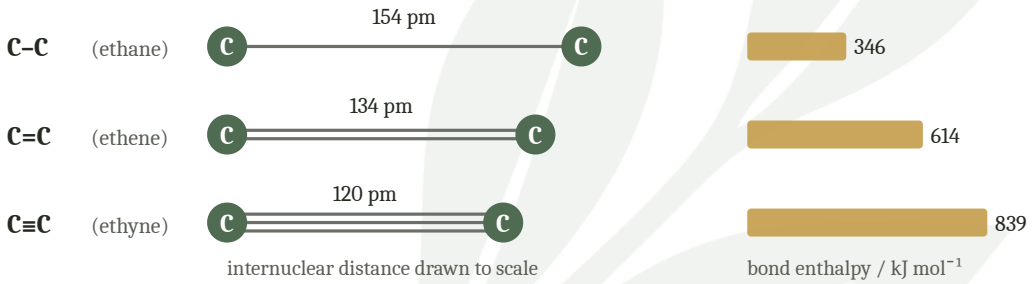


Figure 3 — The three carbon–carbon bonds drawn to scale, with their bond enthalpies (data booklet Sections 11 and 12).

Bond	Shared pairs	Bond length / pm	Bond enthalpy / kJ mol^{-1}	Found in, for example
C–C	1	154	346	ethane, C_2H_6
C=C	2	134	614	ethene, C_2H_4
C≡C	3	120	839	ethyne, C_2H_2
C–O	1	143	358	ethanol; the C–O–H of ethanoic acid
C=O	2	122	804	the C=O of ethanoic acid; CO_2
C≡O	3	113	1077	carbon monoxide (section 4)
N–N	1	146	158	hydrazine, N_2H_4
N=N	2	125	470	N_2F_2, azo compounds
N≡N	3	110	945	nitrogen, N_2

Ethanoic acid, CH_3COOH, is a neat single-molecule illustration. It contains a C–O single bond and a C=O double bond to the same carbon atom, and the double bond is clearly the shorter of the two.

3.4 Bond length down a group — and the fluorine anomaly

Atomic radius increases down a group, so the bonding pair is further from the nuclei in larger atoms. It is attracted less strongly, so bonds get **longer** and usually **weaker** as you go down. The halogens show this clearly — with one striking exception.

Bond	F–F	Cl–Cl	Br–Br	I–I
Bond length / pm	142	199	228	267
Bond enthalpy / kJ mol^{-1}	**159**	242	193	151

The F–F bond is the shortest, yet it is weaker than Cl–Cl and Br–Br. Fluorine atoms are so small that, at such a short bond length, the three lone pairs on each atom are forced very close to those on the other atom. The repulsion between these non-bonding pairs offsets much of the attraction, which weakens the bond. Use this explanation whenever you meet "why is F_2 an anomaly?"

HL **Why a double bond is not twice as strong as a single bond (links to AHL 2.2.15)**

The = and ≡ notation suggests identical bonds stacked together, but they are not identical. A double bond consists of one **sigma (σ) bond**, formed by head-on overlap of orbitals along the bond axis, and one **pi (π) bond**, formed by sideways overlap above and below that axis. A triple bond is one σ bond and two π bonds. Sideways overlap is less effective, so a π bond is weaker than a σ bond. The data show this: C=C is 614 kJ mol^{-1}, less than 2 × 346 = 692; C≡C is 839 kJ mol^{-1}, well short of 3 × 346 = 1038.

3.5 Do multiple bonds make molecules more or less reactive?

This is the Reactivity 2.2 linking question, and the honest answer is: it depends. That is exactly why examiners like it.

- **Nitrogen, N_2:** the N≡N bond (945 kJ mol^{-1}) is one of the strongest in the data booklet. Breaking it needs so much energy that N_2 is very unreactive. Although nitrogen makes up about 78% of the atmosphere and is essential to life, most organisms cannot use N_2 directly.

- **Alkenes, such as ethene:** the C=C bond is stronger *overall* than C–C, yet alkenes are much more reactive than alkanes. A reaction does not need to break the whole double bond. The concentrated region of electron density makes the molecule a target for electron-seeking reagents, and only one of the two bonding pairs has to break. HL That pair is the weaker π bond.

Model answer: "Explain why the carbon–carbon bond in ethyne is shorter than that in ethene." (2 marks)

The C≡C bond in ethyne has three shared pairs of electrons, whereas the C=C bond in ethene has two **[1]**. The greater electron density between the nuclei produces a stronger electrostatic attraction for both nuclei, which pulls them closer together **[1]**.

Examiner traps in 2.2.2

- Stating the trend the wrong way round. More shared pairs means **shorter** and **stronger**. Shorter and stronger always go together for bonds between the same two elements.
- Writing "more bonds so stronger" without explaining *why*: more electrons between the nuclei give greater electrostatic attraction.
- Using bond strength alone to predict reactivity (see section 3.5).
- Comparing lengths of bonds between *different* pairs of elements as if bond order were the only factor. Atomic radius matters too.

4 Structure 2.2.3 — Coordination bonds SL + HL

4.1 One atom supplies both electrons

In every bond so far, each atom contributed one electron to the shared pair. Sometimes, however, **both** electrons come from the same atom: one atom donates a lone pair into a vacancy on the other. This is a **coordination bond**, also called a dative covalent bond. Two conditions must be met:

- the **donor** atom must have a **lone pair** to give, and
- the **acceptor** atom must have a **vacancy** — an empty orbital in its valence shell — to receive it. H^+ has no electrons at all; B in BF_3 and Al in $AlCl_3$ have incomplete octets; transition element ions have empty orbitals.

Definition to learn word-for-word

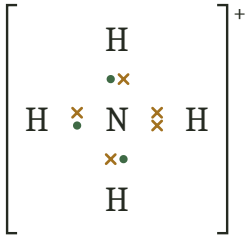
Coordination bond — a covalent bond in which both electrons of the shared pair originate from the same atom.

In line notation, a coordination bond can be drawn as an arrow from the donor to the acceptor (D → A). The arrow records *how* the bond formed. It is not a different type of bond.

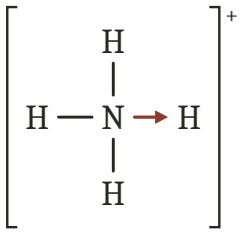
4.2 Once formed, a coordination bond is just a covalent bond

The ammonium ion is the classic example. Ammonia uses the lone pair on nitrogen to bond to a hydrogen ion:

$$NH_3 + H^+ \rightarrow NH_4^+$$



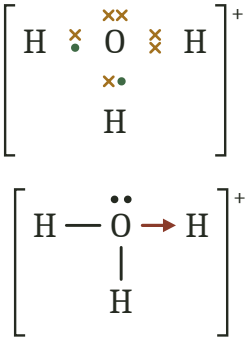
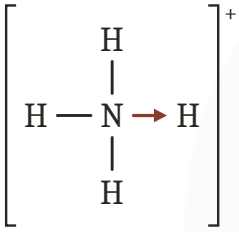
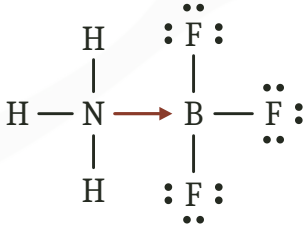
both electrons of the fourth bond came from N (×)

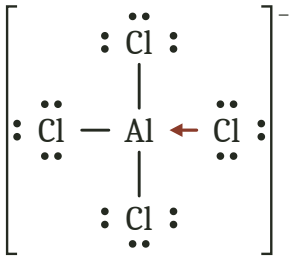
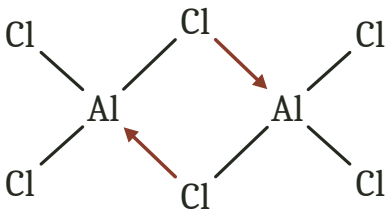


the arrow shows the direction of donation

Once the bond has formed, the electrons are just electrons, and you cannot tell which bond was the coordination bond. All four N–H bonds in NH_4^+ are **identical** in length and strength, and the ion is a perfect tetrahedron. This point is examined directly, so learn the sentence: *a coordination bond, once formed, is indistinguishable from any other covalent bond.*

4.3 The examples you must know

Species	Lewis formula	Donor → acceptor, and what to notice
Hydronium (oxonium) ion, H_3O^+ $H_2O + H^+ \rightarrow H_3O^+$		O → H^+. Water donates one of its two lone pairs, so oxygen is left with one. This is how acids behave in water: $H^+(aq)$ is really H_3O^+.
Ammonium ion, NH_4^+		N → H^+. All four bonds identical once formed.
Carbon monoxide, CO 4 + 6 = 10 electrons		O → C. Of the three shared pairs, two are ordinary (one electron from each atom) and the third is donated entirely by oxygen. The three bonds are nonetheless identical, and C≡O is the strongest bond in the data booklet (1077 kJ mol^{-1}).
Ammonia–boron trifluoride, H_3NBF_3 $NH_3 + BF_3 \rightarrow H_3NBF_3$		N → B. Boron's incomplete octet (six electrons) is completed by the lone pair from nitrogen. A single species containing both a lone-pair donor and an electron-deficient acceptor: a perfect exam example.

Species	Lewis formula	Donor → acceptor, and what to notice
Tetrachloroaluminate ion, $AlCl_4^-$ $AlCl_3 + Cl^- \rightarrow AlCl_4^-$		Cl^- → Al. The same pattern gives BF_4^- from $BF_3 + F^-$.
Aluminium chloride dimer, Al_2Cl_6	 lone pairs on Cl omitted for clarity	Cl → Al, twice. In the vapour at moderate temperatures, two $AlCl_3$ units join: a chlorine on each unit donates a lone pair to the aluminium of the other. Both Al atoms reach an octet.

4.4 How to identify a coordination bond

The skill the guide asks for is *identify*. These four clues will find almost every coordination bond you meet:

1. **A neutral molecule plus H^+.** H^+ has no electrons, so every bond it forms must be a coordination bond: NH_4^+, H_3O^+, PH_4^+.
2. **An atom with more bonds than its usual number** (table in section 2.4). If N has four bonds, O three, or C three with a lone pair, a pair has been donated.
3. **An atom with an incomplete octet that now has eight** — B or Al with four bonds, as in BF_4^-, $AlCl_4^-$, H_3NBF_3 and Al_2Cl_6.
4. **A transition element ion surrounded by molecules or ions with lone pairs** — a complex (section 4.5).

Model answer: "Describe how the bond between nitrogen and boron forms in H_3NBF_3." (2 marks)

The nitrogen atom donates its lone pair of electrons to the boron atom, forming a coordination (dative covalent) bond in which both shared electrons come from nitrogen **[1]**. Boron in BF_3 has only six electrons in its valence shell, so it can accept the pair, which completes its octet **[1]**.

HL

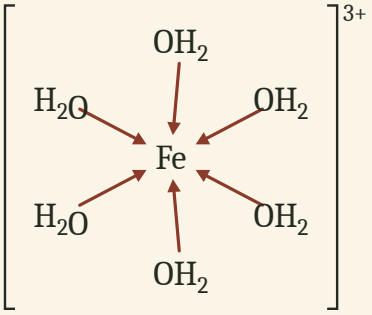
4.5 Lewis acids, Lewis bases and transition element complexes (links to AHL Reactivity 3.4)

The Lewis theory of acids and bases is built on exactly this idea:

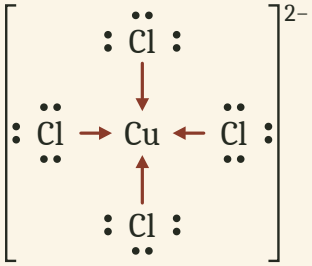
- **Lewis base** — an electron-pair **donor** (NH_3, H_2O, Cl^-, CN^-).
- **Lewis acid** — an electron-pair **acceptor** (H^+, BF_3, $AlCl_3$, transition element ions).

Every reaction between a Lewis acid and a Lewis base forms a coordination bond, because the base donates a lone pair into a vacancy on the acid. This answers the Reactivity 3.4 linking question. The idea is broader than Brønsted–Lowry theory: BF_3 has no hydrogen to donate, yet it is a Lewis acid.

In aqueous solution, transition element ions act as Lewis acids, using empty orbitals in their outer sublevels (for Fe^{3+}, the 4s and 4p) to accept lone pairs. The species that donate are called **ligands**, and the result is a **complex ion**. The number of coordination bonds to the central ion is its **coordination number**.



$[Fe(H_2O)_6]^{3+}$ · six coordination bonds



$[CuCl_4]^{2-}$ · four coordination bonds

When drawing a complex, the arrows always point **from the ligand to the metal ion**, and the bond must come from the atom that carries the lone pair — the O of water, the N of ammonia — not from a hydrogen. The overall charge is the charge on the metal ion plus the charges on any ionic ligands: Cu^{2+} with four Cl^- gives 2+ + 4(1−) = 2−. Lewis acid–base reactions between transition element ions and ligands often produce **coloured** solutions. Adding concentrated hydrochloric acid to pale blue aqueous copper(II) ions, for example, gives yellow-green $[CuCl_4]^{2-}$. Colour is explained in Structure 3.1.

Examiner traps in 2.2.3

- Claiming a coordination bond is weaker, longer or "different" from an ordinary covalent bond once it has formed.
- Pointing the arrow the wrong way. It goes **from** the donor (lone pair) **to** the acceptor.
- Drawing the ligand bond from the H of a water or ammonia ligand instead of from the O or N.
- Saying the donor "gives an electron". It donates a **pair** of electrons.
- Missing the coordination bond in CO because the molecule is neutral and contains no metal.

5 Bringing the three understandings together SL + HL

Bond	Shared pairs	Electrons shared	Where the electrons come from	Length and strength	Examples
Single	1	2	one from each atom	longest, weakest	H–H, C–Cl, O–H
Double	2	4	two from each atom	shorter, stronger	O=O, C=O, C=C
Triple	3	6	three from each atom (except CO)	shortest, strongest	N≡N, C≡C, C≡N
Coordination	1 per bond	2	**both** from one atom	identical to an ordinary bond of the same type once formed	NH_4^+, H_3O^+, CO, H_3NBF_3

One method, applied to anything

Faced with any formula in Paper 1 or Paper 2, work through the same sequence. (1) **Count** the valence electrons, adjusting for charge. (2) **Skeleton**, using the usual-bonds table. (3) **Lone pairs** to the outer atoms first. (4) **Fix** any atom still short by converting lone pairs into multiple bonds, unless it is Be, B or Al, which are allowed incomplete octets. (5) **Inspect**: is any atom carrying more bonds than its usual number? If so, a coordination bond is present. (6) **Brackets and charge** for ions.

Why all this matters for Booklet 2

The Lewis formula is the input to everything that follows. In 2.2.4, the number of electron domains around the central atom, counting lone pairs and treating a multiple bond as one domain, decides the shape of the molecule. In 2.2.5 and 2.2.6, shape and electronegativity decide polarity. Polarity then decides the intermolecular forces and physical properties in 2.2.8 and 2.2.9. A wrong Lewis formula at the start makes every later step wrong.

6 Key points to memorise

Everything below is either wording that examiners look for, or knowledge that is *not* in the data booklet. If you learn nothing else from this booklet, learn this section.

A Definitions — near-verbatim

- **Covalent bond:** the electrostatic attraction between a shared pair of electrons and the positively charged nuclei.
- **Octet rule:** the tendency of atoms to gain a valence shell with a total of eight electrons.
- **Lewis formula:** a diagram showing all valence electrons — bonding pairs and lone (non-bonding) pairs.
- **Lone pair:** a pair of valence electrons not involved in bonding.
- **Single / double / triple bond:** one / two / three shared pairs of electrons.
- **Bond length:** the distance between the nuclei of two bonded atoms.
- **Bond enthalpy:** the energy needed to break one mole of a bond in gaseous molecules (an average value for bonds found in many compounds).
- **Coordination bond:** a covalent bond in which both electrons of the shared pair originate from the same atom.
- HL **Lewis acid / Lewis base:** electron-pair acceptor / electron-pair donor. **Ligand:** a species that donates a lone pair to a central metal ion.

B Usual bonds and lone pairs

Atom	H	C, Si	N, P	O, S	F, Cl, Br, I	Be	B, Al
Bonds	1	4	3	2	1	2	3
Lone pairs	0	0	1	2	3	0	0
Electrons round it	2	8	8	8	8	**4**	**6**

Bold = incomplete octet. An atom with *more* bonds than this table predicts has usually received a donated pair.

C The Lewis method in four words

Count → Skeleton → Complete → Check. Count valence electrons (± charge); join outer atoms to the central atom; give outer atoms lone pairs first; check the total, make multiple bonds if the central atom is short, and bracket ions with the charge outside.

D Numbers worth carrying in your head

- C–C 154 pm / 346 · C=C 134 pm / 614 · C≡C 120 pm / 839 (kJ mol^{-1}).
- N≡N 945 kJ mol^{-1} — the reason N_2 is unreactive. C≡O 1077 kJ mol^{-1} — the strongest bond in the booklet.
- H–H: 74 pm, 436 kJ mol^{-1}. F–F (159) is **weaker** than Cl–Cl (242): lone-pair repulsion in the small F_2 molecule.
- HL C=C < 2 × C–C because a π bond is weaker than a σ bond.

E Coordination bonds in one glance

- Donor needs a **lone pair**; acceptor needs a **vacancy** (H^+, B, Al, transition element ions).
- Must-know examples: NH_4^+, H_3O^+, CO, H_3NBF_3, $AlCl_4^-$, Al_2Cl_6; HL $[Fe(H_2O)_6]^{3+}$, $[CuCl_4]^{2-}$.
- The arrow points donor → acceptor. Once formed, the bond is identical to any other covalent bond.

F Command terms — what each one is actually asking for

- **Deduce** (a Lewis formula) — reach the structure by reasoning from electron counts. Show every lone pair.
- **Identify** (a coordination bond) — point to it: name the bond, or mark it with an arrow. No explanation is needed unless asked.
- **Explain** (length / strength) — give the cause: number of shared pairs → electron density between nuclei → electrostatic attraction.
- **Outline** (how a coordination bond forms) — a brief account: which atom donates which pair to which acceptor.
- **Compare** — mention **both** bonds in every point you make.

7 Check yourself

Cover the answers on the next pages. Marks are shown in brackets where a question is written in exam style. Questions marked HL are for Higher Level students.

1. Define a covalent bond. **[2]**
2. Calculate the total number of valence electrons in: (a) PCl_3 (b) HCN (c) CH_3OH (d) SO_4^{2-} (e) H_3O^+ (f) CN^-
3. Deduce the Lewis formula for each of the following: (a) F_2 (b) H_2S (c) PF_3 (d) OF_2 (e) N_2H_4 (f) CH_3OH
4. Deduce the Lewis formula for each of the following, each of which contains at least one multiple bond: (a) CS_2 (b) CN^- (c) $COCl_2$ (carbon is the central atom)
5. State two different ways in which a species can fail to obey the octet rule, giving one example of each. **[2]**
6. Which molecule contains the shortest carbon–carbon bond?
   A C_2H_6 B C_2H_4 C C_2H_2 D C_2H_5OH
7. Which correctly compares bond enthalpy and bond length for the bonds between two nitrogen atoms?
   A N≡N strongest and longest B N≡N strongest and shortest C N–N strongest and shortest
   D N–N weakest and shortest
8. Explain why the bond enthalpy of N≡N (945 kJ mol^{-1}) is much greater than that of N–N (158 kJ mol^{-1}). **[2]**
9. The H–X bond lengths of the hydrogen halides are 92, 128, 141 and 160 pm. Assign each value to HF, HCl, HBr and HI, and explain the trend. **[2]**
10. Explain why the F–F bond is weaker than the Cl–Cl bond, even though it is shorter. **[2]**
11. Which species contain a coordination bond? I. NH_4^+ II. CO III. CH_4
    A I and II only B I and III only C II and III only D I, II and III
12. Phosphine, PH_3, reacts with a hydrogen ion to form the phosphonium ion, PH_4^+. Draw the Lewis formula of PH_4^+, showing the coordination bond with an arrow, and state how the bond lengths in PH_4^+ compare with one another. **[3]**
13. Aluminium chloride exists in the vapour as Al_2Cl_6 molecules. Explain why two $AlCl_3$ units join together, identifying the type of bond that forms between them. **[3]**
14. Explain why the noble gases rarely form covalent bonds. **[2]**
15. HL In the complex ion $[Co(NH_3)_6]^{3+}$, identify the Lewis acid, the Lewis base and the coordination number, and state the atom in each ligand that forms the bond to cobalt. **[3]**
16. HL Using data from Section 12 of the data booklet, show that the C=C bond is not simply two C–C bonds, and explain why. **[2]**

8 Answers

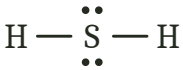
1. The electrostatic attraction **[1]** between a shared pair of electrons and the positively charged nuclei of the two bonded atoms **[1]**.

2. (a) 5 + (3 × 7) = 26 (b) 1 + 4 + 5 = 10 (c) 4 + (4 × 1) + 6 = 14 (d) 6 + (4 × 6) + 2 = 32 (e) 6 + (3 × 1) − 1 = 8 (f) 4 + 5 + 1 = 10

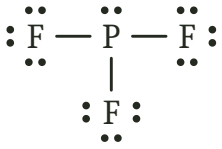
3. Every lone pair must be shown. Either notation is acceptable.



(a) F_2 · 14 e



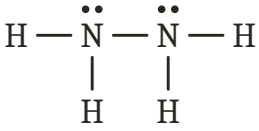
(b) H_2S · 8 e



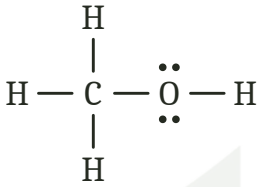
(c) PF_3 · 26 e



(d) OF_2 · 20 e

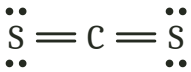


(e) N_2H_4 · 14 e

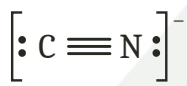


(f) CH_3OH · 14 e

4. Each species contains at least one multiple bond; every lone pair is shown.



(a) CS_2 · 16 e · two double bonds



(b) CN^- · 10 e · triple bond



(c) $COCl_2$ · 24 e · C=O double bond

5. Any two of: an **incomplete octet**, e.g. BF_3, $BeCl_2$, $AlCl_3$ (or H with only two electrons) **[1]**; an **odd number of electrons (radical)**, e.g. NO **[1]**; an **expanded octet**, e.g. PCl_5, SF_6 **[1]**.

6. **C.** Ethyne contains a C≡C triple bond (120 pm). C_2H_4 has C=C (134 pm); C_2H_6 and ethanol have only C–C (154 pm).

7. **B.** More shared pairs give a shorter and stronger bond.

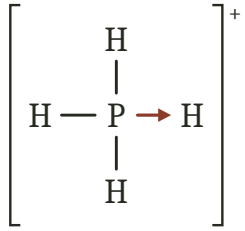
8. N≡N has three shared pairs and N–N only one, so there is a much greater electron density between the nuclei **[1]**. This gives a stronger electrostatic attraction between the shared electrons and both nuclei, so much more energy is needed to break the bond **[1]**.

9. HF 92, HCl 128, HBr 141, HI 160 pm **[1]**. The halogen atom gets larger down the group, so the bonding pair lies further from the halogen nucleus and the internuclear distance increases **[1]**.

10. Fluorine atoms are very small, so in the short F–F bond the lone pairs on the two atoms are very close together **[1]**. The repulsion between these non-bonding pairs weakens the bond, more than offsetting the effect of the shorter distance **[1]**.

11. **A.** NH_4^+ (N → H^+) and CO (O → C). Every bond in CH_4 has one electron from each atom.

12. Lewis formula with four P–H bonds and the charge outside square brackets **[1]**; an arrow from P to one H, showing that phosphorus donated its lone pair **[1]**; all four P–H bonds have the same length, because a coordination bond is identical to the other covalent bonds once formed **[1]**.



13. Aluminium in $AlCl_3$ has only six electrons in its valence shell, an incomplete octet **[1]**. A chlorine atom of one $AlCl_3$ unit donates a lone pair to the aluminium of the other, and vice versa **[1]**, forming two coordination (dative covalent) bonds that give each aluminium an octet **[1]**.

14. Their atoms already have a full valence shell (an octet, or a duet for helium) **[1]**, and they have very high ionization energies, so there is no energetic advantage in sharing electrons **[1]**.

15. Lewis acid: Co^{3+} (electron-pair acceptor) **[1]**; Lewis base: NH_3 (electron-pair donor), bonding through its N atom, which carries the lone pair **[1]**; coordination number 6 **[1]**.

16. $2 \times C–C = 2 \times 346 = 692$ kJ mol^{-1}, but $C=C = 614$ kJ mol^{-1}, which is less **[1]**. A double bond is one σ bond plus one π bond. The π bond, formed by weaker sideways overlap of p orbitals, is weaker than a σ bond **[1]**.

Where this booklet leads

- **Structure 2.2 — Booklet 2**: VSEPR shapes (2.2.4), bond and molecular polarity (2.2.5–2.2.6), covalent network structures (2.2.7), intermolecular forces and properties (2.2.8–2.2.9), chromatography (2.2.10); HL resonance, benzene, expanded octets, formal charge, σ and π bonds, hybridization (2.2.11–2.2.16).
- **Reactivity 1.2** — using the bond enthalpies from section 3.3 to calculate enthalpy changes.
- **Reactivity 3.4** — HL Lewis acids and bases; the coordination bonds of section 4 seen as acid–base reactions.
- **Structure 3.1** — transition element complexes and the origin of their colour.

These notes are a free resource, shared for any IB student who finds them useful.

If a particular idea here still isn't clicking, specialist one-to-one support in Physics, Chemistry and Maths from a past A Level examiner is available at iBLaurel — call or message **+44 7404 710097**, or use the contact form at **www.iBLaurel.com**.

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