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Topic 2.5: Lewis Diagrams

Revision Notes · AP Chemistry

UNIT 2 · COMPOUND STRUCTURE AND PROPERTIES

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What a Lewis Diagram Shows

A Lewis diagram is a map of the valence electrons in a molecule or polyatomic ion. Every valence electron the species owns must appear in the diagram, either in a bonding pair, shared between two atoms, or in a lone pair, held by one atom alone. One line between two atoms is one shared pair, two lines are two shared pairs, and three lines are three shared pairs.

The diagram is a bookkeeping model, not a picture of shape. It fixes connectivity and electron accounting: which atoms are joined, by how many pairs, and how many electrons surround each atom. Almost every judgement in this topic is an audit of that accounting.

Counting the Electrons First

Construction always starts with a single number: the total valence electrons available. Add the valence electrons of every atom in the formula, taking the count from each element's group. Then adjust for charge. Each unit of negative charge adds one electron; each unit of positive charge removes one.

The counts that run through this topic: H_2O has 6 + 1 + 1 = 8. CO_3^{2-} has 4 + (3 × 6) + 2 = 24, where the final 2 comes from the charge. NH_4^+ has 5 + (4 × 1) − 1 = 8. A finished diagram must use exactly its total, no more and no fewer, and most errors in this topic are counting errors made before a single bond is drawn.

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The Construction Sequence

Build the skeleton. The central atom is usually the least electronegative atom that is not hydrogen; hydrogen is always terminal, and halogen atoms are usually terminal as well. Join each terminal atom to the centre with a single bond. Every bond spends 2 of the counted electrons.

Complete the terminal atoms. Distribute the remaining electrons as lone pairs on the outer atoms until each one has an octet. Hydrogen is the exception: it is complete with the 2 electrons of its single bond and never carries a lone pair.

Park what is left on the centre. Any electrons still unspent go on the central atom as lone pairs, even when that takes the centre beyond eight. XeF_4 shows why this step matters: the molecule has 36 electrons in total, the four bonds spend 8 and the four fluorines take 24 as lone pairs, and the last 4 electrons belong on xenon as two lone pairs.

Convert lone pairs into bonds if the centre is short. If the electrons run out while the central atom still has fewer than eight around it, move a lone pair from a terminal atom into the bond between them, making a double bond; repeat, or form a triple bond, until the centre reaches eight. In CO_2, single bonds leave carbon surrounded by only 4 electrons, and converting one lone pair from each oxygen gives O=C=O: every atom has an octet, and all 16 electrons are placed.

The Octet Rule and the Period 2 Ceiling

The octet rule states that atoms in a molecule tend to be surrounded by eight valence electrons, with each shared pair counting toward both atoms it joins. It works because eight electrons fill the s and p subshells of a valence shell. Carbon, nitrogen, oxygen and fluorine follow it in essentially every stable compound.

For period 2 the rule is also a ceiling. The n = 2 shell holds at most eight electrons, so nitrogen, oxygen and fluorine can never be surrounded by more than eight, in any diagram, however the bonds are arranged. A nitrogen atom drawn with five bonds is impossible, not merely unusual.

Incomplete Octets: Too Few Electrons to Go Round

Some central atoms do not own enough electrons for an octet, and their correct diagrams stop short of one. Beryllium in $BeCl_2$ is surrounded by only 4 electrons, in two single bonds. Boron is the standard case: in BF_3 and BH_3, the boron atom has three single bonds, 6 electrons around it, and no lone pairs, and the diagram is finished.

The audit proves it. BF_3 owns 24 electrons; the three bonds use 6 and the fluorine lone pairs use 18, so all 24 are spent and nothing remains for boron. That shortage is chemically real, not a

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drawing failure: electron deficient boron compounds readily accept a lone pair from another species, which is why they behave as Lewis acids.

Expanded Octets: Period 3 and Beyond

A central atom in period 3 or below can be surrounded by more than eight electrons. Phosphorus in PCl_5 has 10 around it, sulfur in SF_6 has 12, and XeF_4 and ICl_4^- each place 12 around the central atom: eight in four bonds plus two lone pairs. Expansion belongs only to central atoms in period 3 and beyond; a terminal oxygen or fluorine never exceeds eight, and neither does any period 2 atom anywhere in the diagram.

Expansion can also be created by adding multiple bonds around a period 3 centre. SO_4^{2-} uses 32 electrons in total. Drawn with four single bonds, sulfur is surrounded by 8; drawn with two double bonds and two single bonds, sulfur is surrounded by 12, and the diagram still uses exactly 32 electrons. Both diagrams are legal Lewis diagrams; choosing which one represents the ion better needs formal charge, which is the subject of Topic 2.6.

Traps That Cost Marks in Topic 2.5

Forgetting the charge in the count. CO_3^{2-} has 24 electrons, not 22, and NH_4^+ has 8, not 9. The charge adjustment is part of the total, and skipping it corrupts every step that follows.

Letting the placed electrons drift from the counted total. A finished diagram must contain exactly the electrons you counted. Dropping XeF_4's final four electrons, or quietly adding a lone pair to the boron in BF_3 (which would place 26 electrons where only 24 exist), both fail the audit.

Expanding a period 2 octet. Nitrogen and oxygen cannot hold ten or twelve electrons. When a diagram seems to demand a fifth bond on nitrogen, the error is upstream: recount the total and the skeleton.

Putting lone pairs on hydrogen. Hydrogen is complete with 2 electrons and never carries a lone pair. Electrons left over after the skeleton belong on the central atom, not parked on hydrogens or added to terminal atoms that already have octets.

Miscounting a multiple bond. A double bond puts 4 electrons around each joined atom, and a triple bond 6. Carbon in O=C=O is surrounded by 8 electrons, not 4, and each nitrogen in N≡N has three bonding pairs plus one lone pair.

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Key Relationships to Remember

- Lewis diagram (SAP-4.A.1): every valence electron shown exactly once, as a bonding pair (a line) or a lone pair on one atom; the placed total must equal the counted total
- Total count: sum of group valence electrons, plus one per unit of negative charge, minus one per unit of positive charge (CO_3^{2-} 24, NH_4^+ 8, XeF_4 36, ICl_4^- 36)
- Construction order: skeleton with single bonds, complete the terminal octets, leftover electrons onto the central atom, then convert lone pairs into multiple bonds until the centre reaches an octet (CO_2: 16 electrons, O=C=O)
- Octet rule: eight electrons around an atom, shared pairs counting for both atoms; hydrogen is complete at 2
- Period 2 ceiling: N, O and F can never exceed eight electrons; expansion is reserved for central atoms in period 3 and below (PCl_5 10, SF_6 12, XeF_4 12)
- Incomplete octets: $BeCl_2$ leaves beryllium with 4 electrons and BF_3 and BH_3 leave boron with 6, because the counted total runs out; that deficiency is what makes boron compounds Lewis acids

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Topic 2.5 Worksheet

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