



## Year 12 / IB1 Intensive Revision Course – Chemistry

This course is aimed at providing students with an opportunity to thoroughly revise their 1st-year IB topics during fully-customizable tutorial sessions, according to the request of students and/or parents. The course will also allow students to "catch up and patch up" – address misunderstandings that they may have accrued during their first year of IB studies.

Students will be pre-evaluated based on their performance on a customized multiple-choice test, with a mix of adapted IB past exam questions and in-house questions. These questions are exclusively based on topics that students should have covered at school. Samples of the diagnostic tests can be found on our [Resources](#) page on [www.akademiatuition.com](http://www.akademiatuition.com).

A standard course will feature 8 to 12 hours of revisions on typical first year school topics, with emphases on the following points:

- Theoretical Chemistry (Topics 2-3):
  - Basic Atomic Structure: Interactions between charged particles – nuclei, electrons, ions, and covalent bonds
  - Theoretical Energetics: light energy, electrostatic potential energy, kinetic energy, and heat energy
- Bonding Theories (Topic 4) – getting towards reality – rationales behind known facts. From shells and octet rule, to orbitals and hybridization. From ionic bonding to covalent bonding to the grey areas.
- Practical Chemistry (Topics 1 and 5–10):
  - Mole and stoichiometry problems
  - Using bonding ideas to help understand different aspects of reactions
    - Physical chemistry – practical energetics, rates and equilibrium
    - Acid–base (proton–transfer) chemistry and redox (electron–transfer) chemistry
    - Organic chemistry

## Sample Course excerpt: drawing Lewis structures

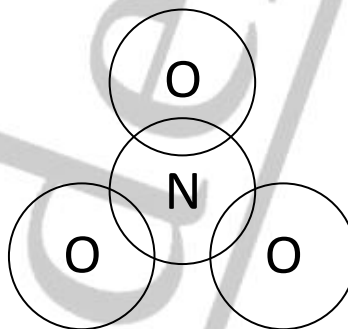
Lewis dot structures (as opposed to IGCSE dot-and-cross structures)

The method outlined below is an extension on the IGCSE or IB-MYP examples, and will apply to a wider range of covalent species. All bonding electrons are "pooled together" to begin with. We represent all electrons with the same dot symbol. We do not care about the origin of these electrons. As a result, we do not have to indicate the presence of dative covalent bonds in these exercises. Using the nitrate ion ( $\text{NO}_3^-$ ) as an example.....

1. Count the number of bonding electrons:

N	5
O x 3	6 x 3 = 18
"minus"	+1
Total	24

2. Construct the atom framework. Usually the atom(s) that need to form more bonds will go into the middle, while atoms that only form 1 bond (like H and F) always stay on the periphery (outside).

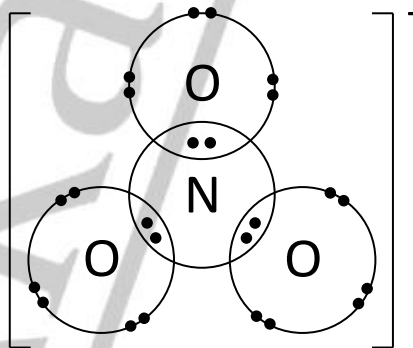


3. Start filling by letting the peripheral atoms (oxygen in this case) satisfy the noble gas configuration.

Are there electrons left?

Yes -- place them as lone pair(s) on central atom

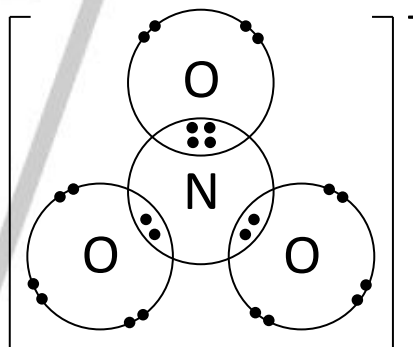
No (this case, all 24 electrons are accounted for) -- next step



4. Do(es) the middle atom(s) have enough electrons?

Yes -- finished

No (this case, nitrogen only has 6 on the previous diagram)-- bring in lone pair(s) from peripheral atoms to form multiple bonds



**Sample Glossary Summary Table**

Topic	Term	IB definition or description as required by the syllabus	Akademia Experts' take
3.1	Lanthanoids	Elements 58 to 71 of the periodic table. These elements have their highest-energy electron occupying the 4f sublevel.	
3.2	Periodicity		Repeating pattern of atomic physical or chemical properties over periods within the periodic table
3.2	Shielding		Repulsions by inner shell electrons of an atom prevent the outermost shell electrons from pulled too close to the nucleus and effectively experiencing the full electrostatic attraction due to all the positive charges within an atomic nucleus.
3.2	Amphoteric substances		Ability of a substance (e.g. $\text{Al}_2\text{O}_3$ ) to behave as a basic or an acidic oxide, depending on what it is being reacted with.
3.2	Electron affinity	The energy change when <u>1 mole of atoms</u> in the <u>gaseous</u> state combine with <u>1 mole of electrons</u> to form 1 mole of 1- anions: $\text{M}(\text{g}) + \text{e}^- \rightarrow \text{M}^-(\text{g})$	Electronegativity values are defined on arbitrary scales that are established by famous scientists, while electron affinity values are measurable energy values. In bonding, the difference in electronegativity values between the two atoms involved are reasonably good guidelines to determine the polarity of a bond, but lots of exceptions exist.
3.2	Electronegativity	The ability of an atom to attract a <u>bonding pair</u> of electrons when involved in a <u>covalent</u> bond	
3.2	First ionization energy	The amount of energy required to <u>remove one mole of electrons</u> from one mole of <u>gaseous atoms</u> , i.e. $\text{M}(\text{g}) \rightarrow \text{M}^+(\text{g}) + \text{e}^-$	
13.1	Explaining catalytic properties of transition metal compounds		Transition metal compounds can behave as catalysts because they can undergo temporary changes in oxidation states and/or temporary changes in geometries during a catalytic cycle.
13.1	Explaining the formation of complex ions	Transition metal atoms and ions have access to vacant (higher-energy) orbitals which can accept electron pairs, when small molecules or ions, called ligands, form dative covalent bonds with them.	Realistically, all metals can do this. For example, $\text{Li}^+(\text{aq})$ is formally $[\text{Li}(\text{H}_2\text{O})_4]^+$ .
13.1	Explaining variable oxidation states of transition metals	The differences in successive ionization energies of a transition metal are similar, and are comparably offset by the increase in the strength of ion-related bonds (such as lattice enthalpy or solution enthalpy) formed by ions of successively higher charges. Therefore ions of transition metals often have similar stabilities.	In contrast, the group 1 and group 2 metals have a clear-cut favorite of +1 and +2 as their most stable oxidation state, respectively.