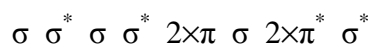


- The electronic energies of the molecular orbitals of diatomics consisting of atoms from H to Ne can be ordered as follows (with energy increasing from left to right):



(the '2×' denotes a pair of degenerate orbitals)

Use this ordering of the molecular orbitals to identify the following species.

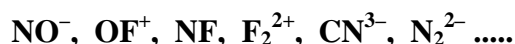
(i) The lowest molecular weight diatomic ion (homo- or heteronuclear) that has **all** of the following characteristics:

- a single negative charge,
- a bond order greater than zero *and*
- is diamagnetic.

HBe⁻ has 6 electrons (1 from He, 4 from Be and 1 from the negative charge) so has a configuration $\sigma^2 \sigma^{*2} \sigma^2$. It has a bond order of 1 and is diamagnetic.

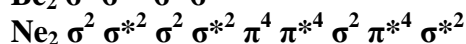
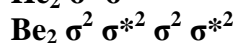
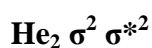
(ii) A diatomic species that has the same electronic configuration as O₂.

There are many: simply substitute one or both O by a cation with the same number of electrons (F⁺, Ne²⁺ etc) or an anion with the same number of electrons (N⁻, C²⁻ etc). For example:



(iii) All of the atoms with atomic numbers less than or equal to 10 that cannot form stable, neutral, homonuclear diatomic molecules.

The neutral, homonuclear diatomic molecule would have a bond order of zero: He, Be and Ne.



Given that there are three degenerate *p* orbitals in an atom, why are there only two degenerate π orbitals in a diatomic molecule?

One *p*-orbital on each atom overlaps end-on with the matching *p*-orbital on the other atom. This produces a σ -bond.

This leaves only two *p*-orbitals on each atom to overlap in the side-on manner required for π bonding.