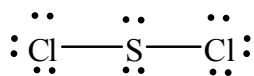
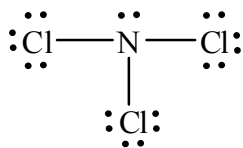


- Draw the Lewis structure for sulfur dichloride, SCl_2 .

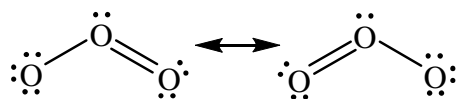
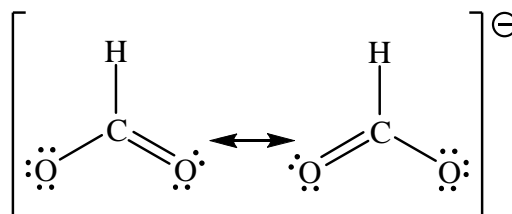
2

- Draw the Lewis structure for nitrogen trichloride, NCl_3 .

2



- Draw Lewis structures of ozone, O_3 , and the formate anion, HCO_2^- , including resonance hybrids where appropriate.

Marks
3 O_3  HCO_2^- 

Marks
1

- Explain what is meant by hybridisation of atomic orbitals.

The mathematical combining of atomic orbitals to produce equivalent orbitals which point in the direction of the bonds and lone pairs.

There is no energy gain in this process, but the overlap of the hybridised orbitals to form bonds results in a lower energy situation because of reduced repulsion of the electron pairs.

- Carbon has atomic number $Z = 6$. What is the ground state electron configuration for an atom of carbon?

5

What compound would you expect to form between a carbon atom with that electron configuration and hydrogen, *i.e.* what is the value of x in the formula CH_x ? Explain.



The bonds would form using the unpaired electrons in the p orbitals. These orbitals would overlap with the $1s$ orbitals of the H atoms.

As there are only 2 unpaired electrons, only 2 bonds would form.

What shape would that molecule have? Explain.

The molecule would be bent with approx 90° bond angle as the p orbitals are at right angles to each other.

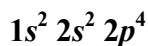
What molecule forms instead? Explain.



The s orbital and the three $2p$ orbitals of carbon undergo hybridisation to form 4 equivalent sp^3 orbitals that point to the corners of a tetrahedron.

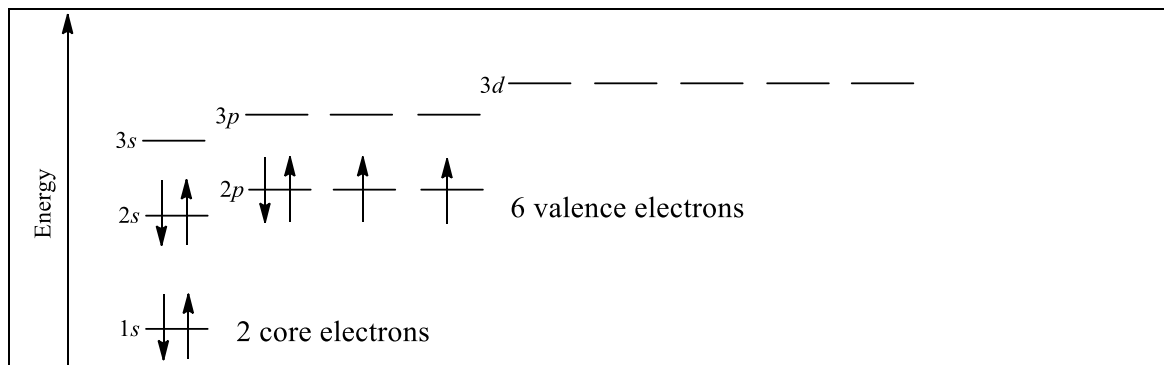
Overlap of these orbitals with the $1s$ orbital of each H atom results in the formation of the tetrahedral molecule, methane.

- What is the ground state electron configuration of oxygen?



Marks
8

The following diagram represents the relative energies of the atomic orbitals in the first three shells. Using arrows to represent electrons, show the most stable electron arrangement of the oxygen atom. Label the core electrons and the valence electrons.



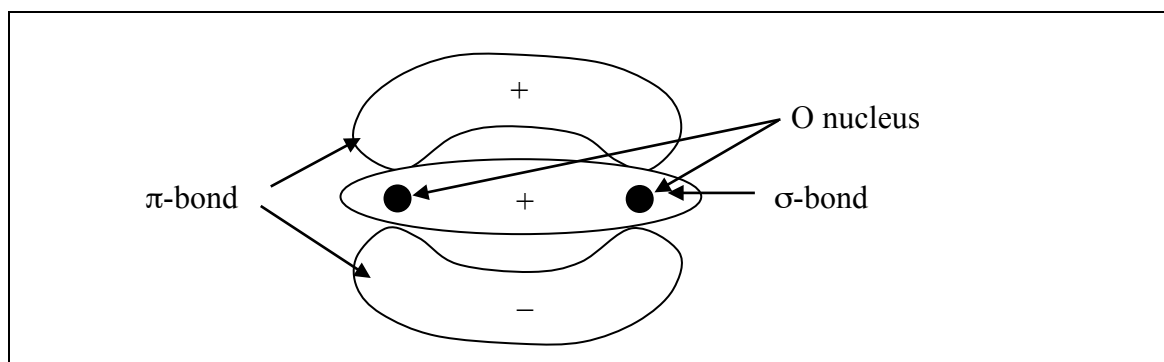
Briefly explain how your diagram illustrates the Pauli exclusion principle, Aufbau principle and Hund's rule.

Pauli exclusion principle: there's a maximum of 2 electrons in each orbital with opposite spins, ensuring that no two electrons have the same set of quantum numbers.

Aufbau principle: lowest energy orbitals fill first.

Hund's rule: electrons in degenerate orbitals (*i.e.* orbitals with same energy) have the maximum number of parallel spins to minimise electron / electron repulsion.

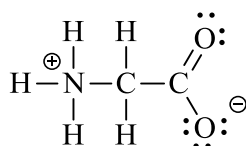
Draw an oxygen molecule showing the shapes of the σ -orbital and the π -orbital present.



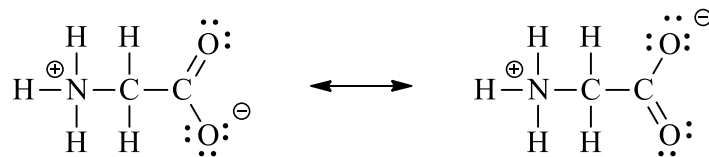
Oxygen and sulfur are both Group 16 elements with a valence of two. Oxygen is a diatomic molecule at room temperature, whilst the bonding in solid sulfur consists only of σ -bonds. Suggest reasons why, at room temperature, the O=O molecule is stable and the S=S molecule is not.

Sulfur would use 3p orbitals to form a π -bond. These orbitals are diffuse and overlap is poor and so it is more stable to use σ -bonds to 2 other atoms. Good overlap of the 2p orbitals in oxygen means that the π -bond is stable.

- Glycine, $\text{NH}_2\text{CH}_2\text{COOH}$, the simplest of all naturally occurring amino acids, has a melting point of $292\text{ }^\circ\text{C}$. The $\text{p}K_a$ of the acid group is 2.35 and the $\text{p}K_a$ associated with the amino group is 9.78. Draw a Lewis structure that indicates the charges on the molecule at the physiological pH of 7.4.



Use your structure to illustrate the concept of resonance.



Describe the hybridisation of the two carbon atoms and the nitrogen atom in glycine and the molecular geometry of the atoms surrounding these three atoms.

N: sp^3 hybridised; tetrahedral geometry

CH₂: sp^3 hybridised; tetrahedral geometry

CO₂⁻: sp^2 hybridised; trigonal planar geometry