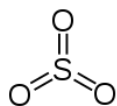


HL Questions on Further aspects of covalent bonding

1. The bonding in sulfur trioxide is sometimes explained by sulfur expanding its octet to six pairs of outer electrons (by using available d orbitals) to give three S=O double bonds.



It is also sometimes explained by keeping to the 'octet rule' where only eight outer electrons around the central sulfur atom are involved. Draw the resonance hybrids to illustrate this model of bonding and explain why all the S–O bonds are the same length and strength.

2. The following is a Lewis structure for carbon dioxide: $\text{I} \text{O} \equiv \text{C} - \text{O} \text{I}$
Show that it has a formal charge of zero and explain why it is not the preferred structure.
3. Explain why the two C–O bond lengths in propanoic acid, $\text{C}_2\text{H}_5\text{COOH}$, are different and yet the two C–O bond lengths in the propanoate anion, $\text{C}_2\text{H}_5\text{COO}^-$, are the same length.
4. Explain why p orbitals on one atom can form two pi bonds and one sigma bond when they combine with the p orbitals on another atom.
5. Ultraviolet light can break the oxygen to oxygen bond in both oxygen, O_2 , and ozone, O_3 , molecules.
The O=O bond energy is 498 kJ mol^{-1} . Calculate the wavelength of light required to break this bond.
Suggest why the light required to break the O–O bond in oxygen is of a higher frequency than the light required to break the O–O bond in ozone.
6. When there is a large amount of delocalised electrons in a molecule the molecule tends to be coloured. The two structures below are for the indicator phenolphthalein in (a) acid solution and (b) in alkali solution. Explain why phenolphthalein becomes coloured in an alkaline solution.

