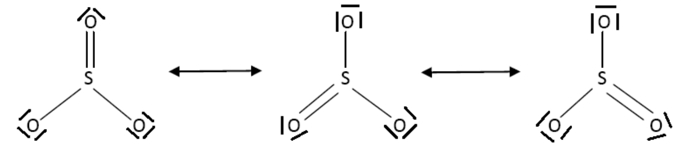
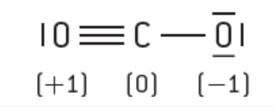
HL Answers to Further aspects of covalent bonding questions

1.



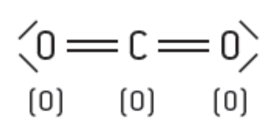
The S–O bonds are the same length and strength as the real structure lies between these three extreme resonance structures with identical S–O bonds with an average of one and a third bonds each (a bond order of 4/3).

2.



C = 4 – 0 – (½ x 8) = 0, O = 6 – 2 – (½ x 6) = +1 (O with triple bond)
O = 6 – 6 – (½ x 2) = – 1(O with single bond) so overall charge = 0

However in the preferred structure with two double bonds all the atoms within the molecule each have a charge of zero.



3. In propanoic acid, C_2H_5COOH, there is a C–O single bond and a C=O double bond. The C–O single bond is longer than the C=O double bond. In the propanoate ion delocalisation occurs so that the extra electron is shared between the two C–O bonds giving two identical bonds with an average 1.5 bond which has a bond length between a C–O single bond and a C=O double bond.

4. Two of the p orbitals combine 'sideways' to form pi bonds whereas the other p orbital combines 'head on' forming a sigma bond.

5. For just one double bond this equates to 498 divided by Avogadro's constant = 8.27×10^{-19} J. The wavelength of light that corresponds to this enthalpy value (E) is calculated by combining the expressions $E = h\nu$ and $c = \lambda\nu$ to give $\lambda = \frac{hc}{E}$

$$\lambda = \frac{6.63 \times 10^{-34}\,(\text{Js}) \times 3.00 \times 10^{8}\,(\text{ms}^{-1})}{8.27 \times 10^{-19}\,(\text{J})} = 241\ \text{nm}$$

The double O=O bond in oxygen is stronger than the 1.5 O-O bond in ozone (the average of the two resonance hybrid structures) so more energy (higher frequency) is required to break the O-O bond in oxygen.

6. The hybridization of the central carbon atom changes from sp^3 to sp^2. This leaves a single electron in a p orbital so that the electrons can delocalise throughout the whole anion via conjugated π bonds whereas in the undissociated molecule the delocalisation is limited mainly just to the separate aromatic rings.