3.091 OCW Scholar

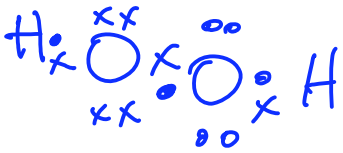
Self-Asessment
Bonding and Molecules

Supplemental Exam Problems for Study
Solutions Key

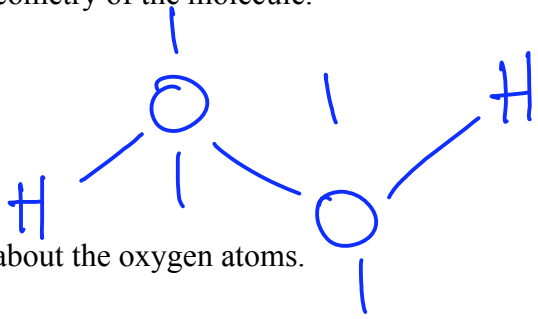
Problem #1

Answer the following questions about hydrogen peroxide (H_2O_2).

(a) Draw the Lewis structure of H_2O_2.



(b) Draw a 3-dimensional representation of the molecular geometry of the molecule.



(c) Name the geometry of the electron distribution about the oxygen atoms.

tetrahedral

(d) Determine the per cent ionic character of the O–H bond.

$\chi_O = 3.44$ $\chi_H = 2.20$ $\therefore \Delta\chi = 1.24 \Rightarrow$ % ionic character is

$[1 - \exp\{-\frac{1}{4}(\Delta\chi)^2\}] \times 100 = 32\%$

(e) Is the molecule polar or nonpolar? Explain.

nonpolar – symmetric dispersal of polar bonds

(f) Is it chiral or achiral? Explain.

achiral – symmetric dispersal of atoms around center of symmetry

(g) Calculate the maximum wavelength of a beam of neutrons capable of breaking the O–H bond in H_2O_2.

DATA: Average Bond Energies (kJ/mol)

O–O	142
H–H	432

$$E_{OH} = \sqrt{E_{O\text{-}O} \cdot E_{H\text{-}H}} + 96.3(\chi_O - \chi_H)^2 = (142 \cdot 432)^{1/2} + 96.3(1.24)^2$$

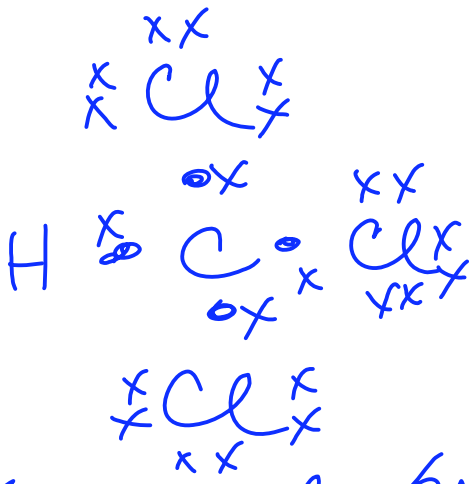
$$= 248 + 148 = 396 \text{ kJ/mol} / N_A = 6.57 \times 10^{-19} \text{ J/bond}$$

$$E_{neutron} = \frac{p^2}{2m} = \frac{h^2}{2m\lambda^2} = E_{bond} \Rightarrow \lambda = \frac{h}{(2mE)^{1/2}}$$

$$\therefore \lambda_{neutron} = \left(\frac{6.6 \times 10^{-34}}{2 \times 1.67 \times 10^{-27} \times 6.57 \times 10^{-19}}\right)^{1/2} = 1.41 \times 10^{-11} \text{ m}$$

Problem #2

(a) Draw the Lewis structure of trichloromethane ($CHCl_3$).



(b) Is $CHCl_3$ polar or nonpolar? Explain.

polar – asymmetric molecule with polar bonds between different atom pairs

(c) Calculate the maximum wavelength of electromagnetic radiation capable of breaking the C–Cl bond in $CHCl_3$.

DATA: bond energy (kJ/mol)

C–C = 346

Cl–Cl = 240

H–H = 432

$E_{C-Cl\ bond} \leq E_{e\text{-}m\ radiation}$

$E_{C-Cl} = \sqrt{E_{CC} \cdot E_{ClCl}} + 96.3(\chi_C - \chi_{Cl})^2$

$= (346 \cdot 240)^{1/2} + 96.3(2.55 - 3.16)^2$

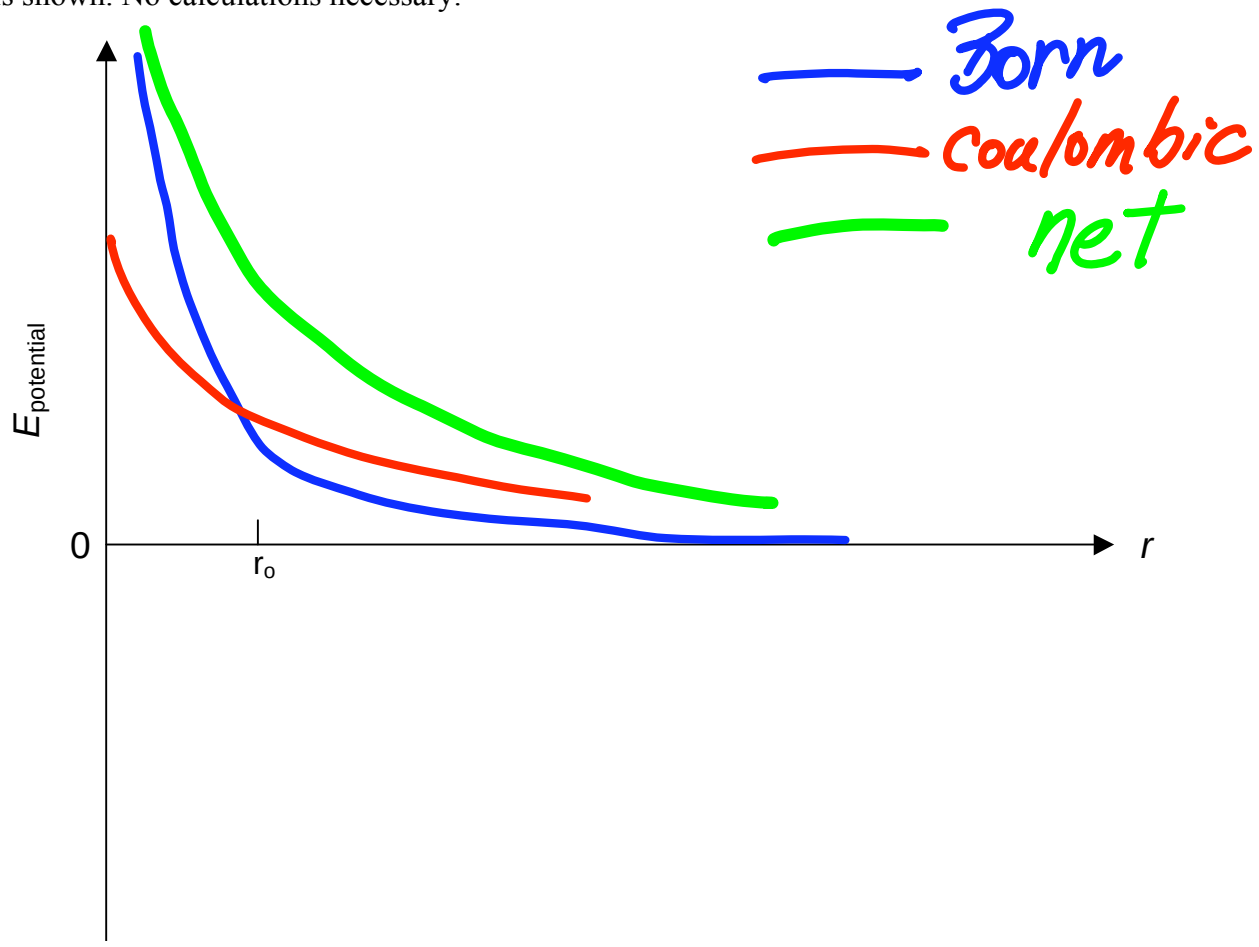
$= 288 + 36 = 324\ kJ/mol = E_{photon} = \frac{hc}{\lambda}$

$\therefore \lambda = \frac{hc}{E_{C-Cl}} = \frac{(6.6 \times 10^{-34})(3 \times 10^8)}{\frac{324 \times 10^3}{6.02 \times 10^{23}}}$

$= 3.68 \times 10^{-7}\ m$

Problem #3

Sketch the relationship between potential energy ($E_{potential}$) and internuclear distance (r) for the interaction between a bromide ion (Br^-) and an iodide ion (I^-). For reference, the distance $r_o = r_{Br^-} + r_{I^-}$ is shown. No calculations necessary.



Problem #4

(a) For each set of chemical species, rank in order of boiling point from lowest to highest. Justify with reference to the operative chemical bonding.

(i) Ar and HCl and F_2

HCl is polar; Ar and F_2 are nonpolar. ∴HCL has the highest b.p.

Ar and F_2 have the same number of e⁻s but Ar is more polarizable – no e⁻s are confined to bonds

∴Ar b.p. is higher than that of F_2

so the rank order of b.p.s is F_2 < Ar < HCl

by the way, you can look up the b.p.s and polarizabilities of Ar and F_2 on the P.T. so there is no mystery here ☺

(ii) CH_4 and CF_4 and HF

HF is polar and has H-bonding capability; CH_4 and CF_4 are nonpolar. ∴HF has the highest b.p.

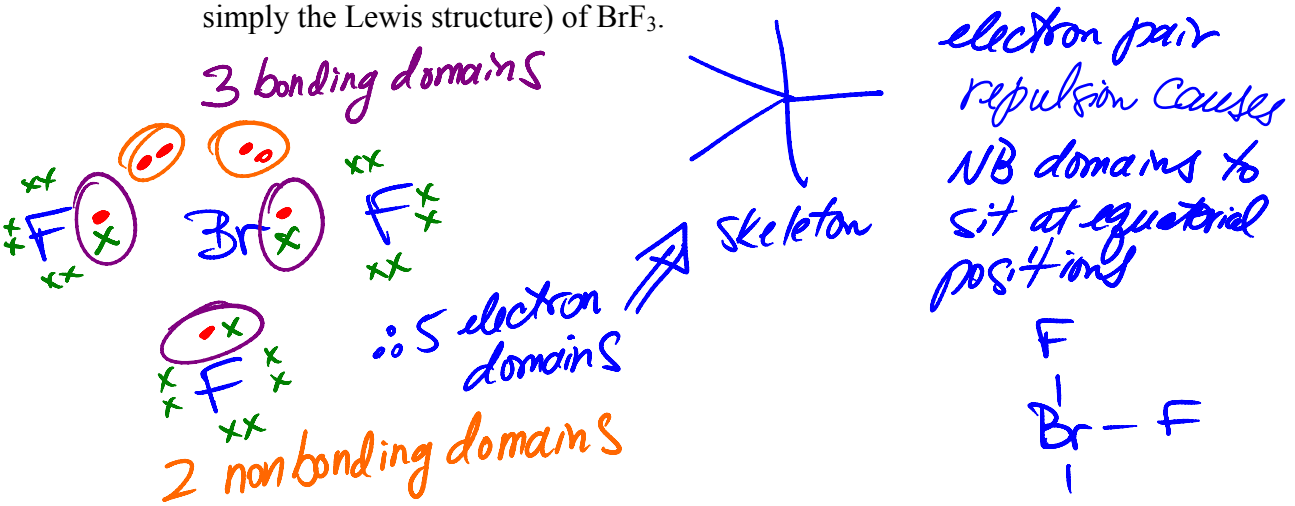
CH_4 and CF_4 have the same molecular structure but CF_4 has more e⁻s and is larger and is therefore more polarizable

∴ CF_4 b.p. is higher than that of CH_4

so the rank order of b.p.s is CH_4 < CF_4 < HF

Problem #5

(a) Draw a 3-dimensional representation of the molecular geometry around the central atom (not simply the Lewis structure) of BrF_3.



(b) Name the type of hybrid orbitals that the central atom forms.

sp^3d

(c) State whether the molecule is polar or nonpolar. Justify.

polar – asymmetric placement of polar bonds around central atom

$\delta+$ Br – F $\delta-$

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