

Chemguide – answers

BOND ENTHALPIES

1. a) It means that it takes 193 kJ to break the bonds in 1 mole of gaseous bromine to produce gaseous bromine atoms. (If you haven't stated clearly that both the original bromine and the atoms formed have to be gaseous, then it is wrong.)
- b) The amount of energy needed to break a bond depends on the environment of the bond. Every time you remove one of the hydrogen atoms from the carbon atom in methane, the environment of what is left changes. So each individual carbon-hydrogen bond needs a different amount of energy to break it. The mean (or average) bond enthalpy is measured by finding out how much energy in total is needed to break a mole of methane into its constituent atoms and then dividing by 4 to give an average value. In the bromine case, this problem doesn't arise because you are only breaking a unique bond in the molecule.
- c) In ethane, the hydrogen is attached to a carbon atom in a carbon-carbon single bond; in ethene, it is attached to a carbon in a carbon-carbon double bond. The environment is different, and so you would expect the bond enthalpy to be slightly different.
2. a) Because you aren't given a value for the bond enthalpy of the C-C bond, you have to do this by the quick method, looking only at which bonds are broken and made.

Energy needed to break C-H and Cl-Cl:

$$+413 + 243 = +656 \text{ kJ mol}^{-1}$$

Energy released when you make C-Cl and H-Cl:

$$-346 - 432 = -778 \text{ kJ mol}^{-1}$$

$$\text{Overall change} = +656 - 778 \text{ kJ mol}^{-1} = -122 \text{ kJ mol}^{-1}$$

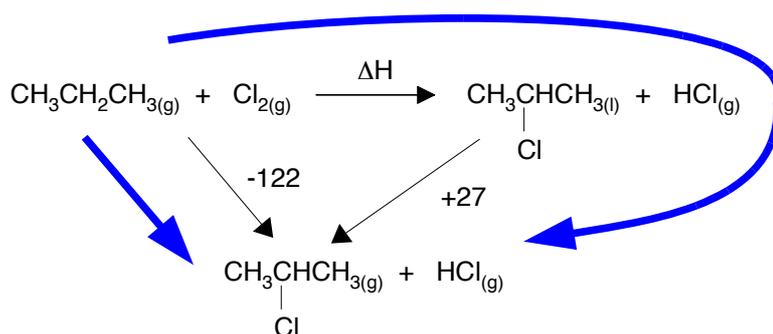
(If you look back at the Chemguide page, you will find that this is exactly the same as the enthalpy change during the chlorination of methane. That is because you are breaking and making the same bonds in both cases. In practice, the values won't be *exactly* the same because some of the bonds are in slightly different environments.)

b) The quickest way of doing this is to realise that when 1 mole of 2-chloropropane gas turns to liquid, the enthalpy change is going to be -27 kJ mol^{-1} . That is the reverse of the enthalpy change of vaporisation.

The bond enthalpy calculation shows that the enthalpy change to get as far as 2-chloropropane gas is -122 kJ mol^{-1} . To go to the liquid, you will get an extra -27 kJ mol^{-1} . The total enthalpy change will be -149 kJ mol^{-1} .

Alternatively, you could use a Hess's Law cycle diagram, putting in the information you already know from part (a). The diagram for this is on the next page. You will need to look very closely at the state symbols.

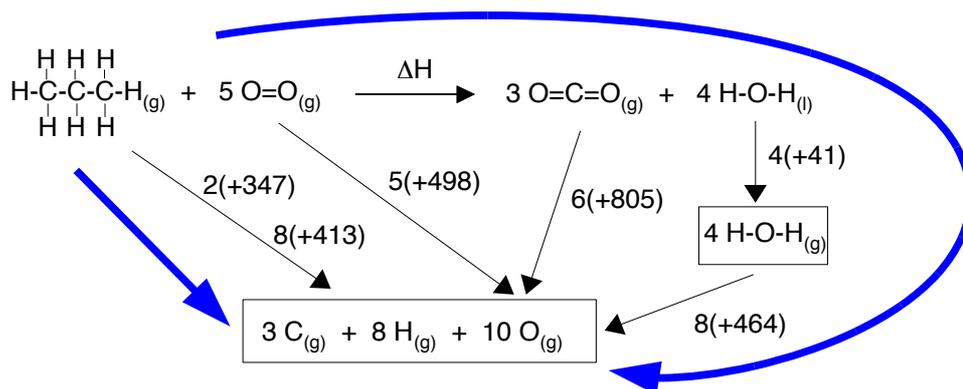
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$$\Delta\text{H} + 27 = -122$$

$$\Delta\text{H} = -149 \text{ kJ mol}^{-1}$$

3. Personally, when it gets this complicated, I find it much easier to just put on the numbers on to a Hess's Law cycle. You are much less likely to get a sign wrong as well, particularly if you are having to use an enthalpy change of vaporisation.



$$\Delta\text{H} + 6(805) + 4(41) + 8(464) = 2(347) + 8(413) + 5(498)$$

$$\Delta\text{H} = 2(347) + 8(413) + 5(498) - 6(805) - 4(41) - 8(464)$$

$$\Delta\text{H} = -2218 \text{ kJ mol}^{-1}$$

If you got this wrong, don't leave it until you understand exactly where the problem lies. The most likely reason is that you haven't counted the bonds properly.

These questions only scratch the surface of bond enthalpy calculations. If you got them right, well done, but you now need to use another source of questions to look at all the variations on this that you might come across.