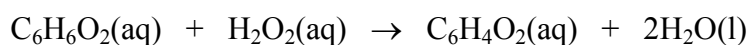
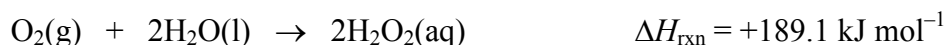
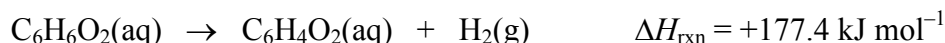


Marks
8

- The conversion of hydroquinone ($C_6H_6O_2(aq)$) to quinone ($C_6H_4O_2(aq)$) is involved in many important biochemical reactions. The bombardier beetle, for example, uses the explosive reaction between hydroquinone and hydrogen peroxide (as described by the equation below) as a defence mechanism.



From the following reaction data, calculate ΔH_{rxn} for the reaction between 1.00 mol of hydroquinone and 1.00 mol of hydrogen peroxide.



$O_2(aq)$ and $H_2O_2(aq)$ and this is related to reaction (1) and the reverse of reaction (2). As the former generates $H_2(g)$ and the latter $O_2(g)$ neither of which are present in the conversion reaction, they are combined with the reverse of reaction (3).

To ensure a balance equation, the combination is thus (1) – $\frac{1}{2} \times$ (2) – (3):

				ΔH_{rxn}
(1)	$C_6H_6O_2$	\rightarrow	$C_6H_4O_2 + H_2$	177.4
– $\frac{1}{2} \times$ (2)	H_2O_2	\rightarrow	$\frac{1}{2}O_2 + H_2O$	– $\frac{1}{2} \times 1891.$
– (3)	$H_2 + \frac{1}{2}O_2$	\rightarrow	H_2O	– 285.8
	$C_6H_6O_2 + H_2O_2$	\rightarrow	$C_6H_4O_2 + 2H_2O$	–203 kJ mol ^{–1}

$$\Delta H_{rxn} = -203 \text{ kJ mol}^{-1}$$

Use the answer you obtained above to calculate the heat liberated (in joules) in the oxidation of 3.86×10^{-4} mol of hydroquinone to quinone.

As 203 kJ are liberated by 1 mol, the heat change for 3.86×10^{-4} mol is:

$$q = (203 \text{ kJ mol}^{-1}) \times (3.86 \times 10^{-4} \text{ mol}) = 0.0784 \text{ kJ} = 78.4 \text{ J}$$

Answer: 78 J

Calculate the temperature rise of 0.250 g of water for this quantity of heat. (The heat capacity of water, $C_p = 4.184 \text{ J K}^{-1} \text{ g}^{-1}$)

Using, $q = m \times C \times \Delta T$,

$$\Delta T = \frac{q}{m \times C} = \frac{(78.4 \text{ J})}{(0.250 \text{ g}) \times (4.184 \text{ J K}^{-1} \text{ g}^{-1})} = 74.9 \text{ K}$$

Answer: 74.9 K