Marks 5

- Some micro-organisms thrive under warm, acidic conditions where sulfuric acid is produced as a metabolic by-product from the reaction between sulfur (S), water and oxygen (O_2). Write a balanced equation for this reaction.

$2S(s) + 3O_2(g) + 2H_2O(l) \rightarrow 2H_2SO_4(l)$

Calculate the volume of oxygen that is required to react to completion with 0.0655 g of sulfur at 1.00 atm and 55 °C.

The number of moles of sulfur in 0.0655 g is:

$$\text{number of moles} = \frac{\text{mass (in g)}}{\text{molar mass (in g mol}^{-1})} = \frac{0.0655}{32.07} = 0.002042\ \text{mol}$$

Three **moles of O_2 are required for every *two* moles of S. Hence the number of moles of O_2 required is:**

$$\text{number of moles} = \frac{3}{2} \times 0.002042 = 0.003064\ \text{mol}$$

The volume of this number of moles of gas at 1.00 atm and 55 °C (= 328 K) is given by the ideal gas law, pV = nRT, so,

$$V = \frac{nRT}{p} = \frac{(0.003064) \times 0.08206 \times 328}{1.00} = 0.0825\ \text{L}$$

(Note the use of R = 0.08206 L atm K^{-1} mol^{-1} as pressure is given in atmospheres provides the answer in litres)

ANSWER: **0.0825 L**

Calculate the pH of the final solution if the reaction is carried out in 20.0 L of water. Assume that the sulfuric acid fully dissociates.

Two **moles of H_2SO_4 is produced from every *two* moles of S. Therefore, the number of moles of H_2SO_4 is just 0.002042 mol. Every mole of H_2SO_4 produces *two* moles of H^+. Therefore, the number of moles of H^+ is**

number of moles = 2 × 0.002042 = 0.004084 mol

As the reaction is carried out in 20.0 L of water, the concentration of H^+ is:

$$\text{concentration} = \frac{\text{number of moles (in mol)}}{\text{volume (in L)}} = \frac{0.004084}{20} = 0.0002042\ \text{M}$$

From the definition of pH:

$pH = pH = -\log_{10}[H^+] = -\log_{10}(0.0002042) = 3.69$

ANSWER: **3.69**

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