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:

$$\text{number of moles} = \frac{\text{mass}}{\text{atomic mass}} = \frac{0.0655 \text{ g}}{32.07 \text{ g mol}^{-1}} = 0.00204 \text{ mol}$$

From the chemical equation, *two* moles of S react with *three* moles of O_2. Therefore, 1.5 × 0.00204 mol = 0.00306 mol of oxygen is required.

Using the ideal gas law, $PV = nRT$, the volume of this number of moles of O_2 can be calculated. As the pressure = 1.00 atm, the corresponding value of R to be used is 0.08206 L atm K^{-1} mol^{-1}. T = 55 °C = (55 + 273) K = 328 K.

$$V = \frac{nRT}{P} = \frac{(0.00306 \text{ mol})(0.08206 \text{ L atm K}^{-1} \text{ mol}^{-1})(328 \text{ K})}{1.00 \text{ atm}}$$
$$= 0.0825 \text{ L or } 82.5 \text{ mL}$$

ANSWER: **0.0825 L or 82.5 mL**

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 are produced for every *two* moles of sulfur used. The number of moles of H_2SO_4 produced is therefore 0.00204 mol.**

If the H_2SO_4 completely dissociates into H^+ and SO_4^{2-}, *two* moles of H^+ are produced for every *one* mole of H_2SO_4. The number of moles of H^+ is therefore 0.00408 mol.

This number of moles is dissolved in 20.0 L of water, the concentration of H^+ is therefore:

$$\text{concentration} = \frac{\text{number of moles}}{\text{volume}} = \frac{0.00408 \text{ mol}}{20.0 \text{ L}} = 0.000204 \text{ M}$$

Using the definition of pH,

$$\text{pH} = -\log_{10}[\text{H}^+] = -\log_{10}(0.000204) = 3.69$$

ANSWER: **3.69**