Marks 8

- One of the causes of acid rain is a reaction occurring in the upper atmosphere between gaseous NO_2 and water to produce nitric acid and gaseous NO. Write a balanced chemical equation for this reaction.

$3NO_2(g) + H_2O(l) \rightarrow 2HNO_3(aq) + NO(g)$

As part of their school project on acid rain, some high school students collected a sample of rain (220 mL) and measured the pH value of the solution, reporting the value as pH = 3.9. Assuming that the rain sample does not contain any acids other than nitric acid, calculate the volume of gaseous NO_2 that would have been consumed in the upper atmosphere (where temperature = –56 °C and pressure = 11.6 kPa) to produce the sample of rain collected by the students.

By definition,

$$\mathbf{pH = -\log_{10}[H^+] \quad or \quad [H^+] = 10^{-pH}}$$

pH = 3.9 therefore corresponds to $[H^+] = 10^{-3.9}$ M.

The number of moles of H^+ in the sample of 220 mL is,

$$\text{number of moles} = \text{volume (in L)} \times \text{molarity (in mol L}^{-1}\text{)}$$
$$= 0.220\ \text{L} \times 10^{-3.9}\ \text{M} = 2.8 \times 10^{-5}\ \text{mol}$$

As nitric acid is a strong acid, this is also the number of moles of nitric acid present. From the chemical equation, *two* moles of nitric acid are produced from every *three* moles of NO_2. Therefore, the number of moles of NO_2 used is:

$$\text{number of moles} = \frac{3}{2} \times 2.8 \times 10^{-5}\ \text{mol} = 4.2 \times 10^{-5}\ \text{mol}$$

The volume of this number of moles of gas at 11.6 kPa and –56 °C (= 217 K) is given by the ideal gas law:

$$pV = nRT$$

so,

$$V = \frac{nRT}{P} = \frac{(4.2\times10^{-5}\ \text{mol})(8.314\ \text{J}^{-1}\ \text{K}^{-1}\ \text{mol}^{-1})(217\ \text{K})}{11.6\times10^{3}\ \text{Pa}}$$
$$= 6.5 \times 10^{-6}\ \text{m}^3 = 6.5 \times 10^{-3}\ \text{L}$$

(Note the use of $R = 8.314$ J K^{-1} mol^{-1} as pressure is given in Pascal provides the answer in m^3)

ANSWER: **6.5×10^{-3} L**