Marks 5

- Calcium carbide, CaC_2, reacts with water to produce a gas and a solution containing OH^- ions. A sample of CaC_2 was treated with excess water and the resulting gas was collected in an evacuated 5.00 L glass bulb. At the completion of the reaction, the pressure inside the bulb was 1.00×10^5 Pa at a temperature of 26.8 °C. Calculate the amount (in mol) of the gas produced.

5.00 L corresponds to 5.00×10^{-3} m^3 and 26.8 ^{o}C corresponds to (26.8 + 273.0) K = 299.8 K. Using the ideal gas law:

$$PV = nRT$$

$$n = PV / RT$$
$$= (1.00 \times 10^5 \text{ Pa})(5.00 \times 10^{-3} \text{ m}^3) / ((8.314 \text{ Pa m}^3 \text{ mol}^{-1} \text{ K}^{-1})(299.8 \text{ K}))$$
$$= 0.201 \text{ mol}$$

Answer: **0.201 mol**

Given that the mass of the gas collected was 5.21 g, show that the molar mass of the gas is 25.9 g mol^{-1}.

As the number of moles = mass / molar mass:

molar mass = mass / number of moles = 5.21 g / 0.201 mol = 25.9 g mol^{-1}

Suggest a molecular formula for the gas and write a balanced equation for the reaction that occurred.

As the gas was produced from CaC_2 which contains C_2^{2-}, a likely formula is C_2H_2:

molar mass = (2 × 12.01 (C) + 2 × 1.008 (H)) g mol^{-1} = 26.0 g mol^{-1}

This is formed by addition of H_2O:

$$CaC_2(s) + 2H_2O(l) \rightarrow Ca^{2+}(aq) + 2OH^-(aq) + C_2H_2(g)$$