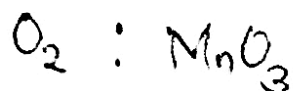


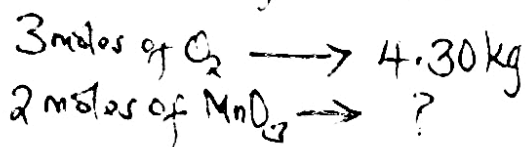
mole ratio for the reaction is 2 : 3 : 2 for Mn : O₂ : MnO₃

If 4.30 kg of oxygen completely reacts with Mn to produce MnO₃, then the mass of MnO₃ produced according to the mole ratios can be calculated as:



mole ratio 3 : 2

3 moles of O₂ reacts to produce 2 moles of MnO₃.
hence mass of MnO₃ produced is given by.



$$\left(\frac{2 \text{ moles} \times 4.30 \text{ kg}}{3 \text{ moles}} \right) = 2.87 \text{ kg of MnO}_3 \text{ produced.}$$

Molar mass

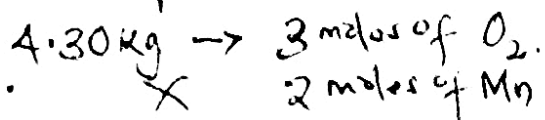
Mn — 54.94 g/mol.

MnO₃ — 102.94 g/mol.

O₂ — 15.999 g/mol.

$$\begin{aligned} \text{Mass of MnO}_3 \\ \text{in g} &= \underline{\underline{2.87 \times 10^3 \text{ g}}} \end{aligned}$$

Mass of Mn that reacted.



$$\left(\frac{4.30 \text{ kg} \times 2 \text{ moles}}{3 \text{ moles}} \right) = \underline{\underline{2.87 \text{ kg of Mn}}}$$

$$\text{Mass in g} = \left(\frac{2.87}{1000} \right) 2.87 \times 10^3 \text{ g of Mn}$$

Mass of MnO₃ in g is the same as mass of Mn since mole ratio is same.

$$\begin{aligned} \text{Moles of MnO}_3 &= \frac{\text{mass of MnO}_3 (\text{g})}{\text{Molar mass (g/mol)}} \\ &= \left(\frac{2.87 \times 10^{-3} \text{g}}{102.94 \text{g/mol}} \right) = \underline{\underline{2.788 \times 10^{-5} \text{ moles of MnO}_3}} \end{aligned}$$

$$\text{mass in g of O}_2 = \left(\frac{4.30}{1000} \right) = \underline{\underline{4.30 \times 10^{-3} \text{g O}_2}}$$

$$\begin{aligned} \text{Moles of Mn} &= \frac{\text{mass of Mn (g)}}{\text{Molar mass (g/mol)}} \\ &= \left(\frac{2.87 \times 10^{-3} \text{g}}{54.94 \text{g/mol}} \right) \\ &= \underline{\underline{5.224 \text{ moles of Mn}}} \end{aligned}$$

$$\text{mass of O}_2 = 4.30 \text{kg}$$

$$\begin{aligned} \text{Moles of O}_2 &= \left(\frac{\text{mass of O}_2}{\text{molar mass (g/mol)}} \right) \\ &= \left(\frac{4.30 \times 10^3 \text{g}}{16 \text{g/mol}} \right) \\ &= \underline{\underline{2.68 \times 10^2 \text{ moles of O}_2}} \end{aligned}$$

MnO₃

Mn : O₂
1 : 3

$$\left(\begin{array}{cc} \boxed{2.87} & \boxed{4.30} \\ \text{Starting Amount.} & \end{array} \right) \left(\begin{array}{cc} \boxed{0.7175} & \boxed{2.1525} \\ \hline \boxed{0.7175} & \boxed{6.4575} \end{array} \right)$$