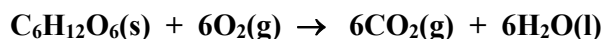


1.



(a) Using  $\Delta_{\text{rxn}}H^\circ = \sum m\Delta_f H^\circ(\text{products}) - \sum n\Delta_f H^\circ(\text{reactants})$ ,

$$\begin{aligned}\Delta_{\text{rxn}}H^\circ &= [6\Delta_f H^\circ(\text{CO}_2(\text{g})) + 6\Delta_f H^\circ(\text{H}_2\text{O}(\text{l}))] - [\Delta_f H^\circ(\text{C}_6\text{H}_{12}\text{O}_6(\text{s})) + 6\Delta_f H^\circ(\text{O}_2(\text{g}))] \\ &= [(6 \times -393) + (6 \times -285)] - [(-1285) + (6 \times 0)] \text{kJ mol}^{-1} \\ &= -2783 \text{kJ mol}^{-1}\end{aligned}$$

As  $\Delta_f H^\circ(\text{O}_2(\text{g})) = 0$  as it is in its standard state.

(b) Using  $\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ$ ,

$$\begin{aligned}\Delta G^\circ &= (-2783 \times 10^3 \text{ J mol}^{-1}) - (298 \text{ K} \times 181 \text{ J K}^{-1} \text{ mol}^{-1}) \\ &= -2837000 \text{ J mol}^{-1} = -2837 \text{ kJ mol}^{-1}\end{aligned}$$

Note that  $\Delta S^\circ$  is expressed in  $\text{J K}^{-1} \text{ mol}^{-1}$  so that  $\Delta H^\circ$  must be converted from  $\text{kJ mol}^{-1}$  into  $\text{J mol}^{-1}$  (by multiplying by  $10^3$ ).

(c) Oxidation of 1 mole of glucose produces  $-2837 \text{ kJ mol}^{-1}$ . If ATP synthesis is 41% efficient then:

$$\text{amount of usable energy} = 0.41 \times (2837 \text{ kJ mol}^{-1}) = 1163 \text{ kJ mol}^{-1}.$$

To produce 1 mole of ADP requires  $30.5 \text{ kJ mol}^{-1}$  so 1200 kJ will produce:

$$\text{number of moles} = \frac{1200 \text{ kJ}}{30.5 \text{ kJ mol}^{-1}} = 39 \text{ mol}$$

(d) The positive sign of  $\Delta G^\circ$  means that the synthesis reaction is not spontaneous. Energy has to be provided to drive the reaction. This energy is supplied by the combustion of glucose so that the overall process (simultaneous synthesis of ATP and combustion of glucose) is spontaneous with  $\Delta G^\circ < 0$ .

2. The molar mass of  $\text{O}_2$  is  $2 \times 16.00 = 32.00 \text{ g mol}^{-1}$ . Therefore, 4.0 g corresponds to:

$$\text{number of moles} = \frac{\text{mass}}{\text{molar mass}} = \frac{4.0 \text{ g}}{32.00 \text{ g mol}^{-1}} = 0.125 \text{ mol}$$

Using the perfect gas law,  $PV = nRT$ :

$$V = \frac{nRT}{P} = \frac{(0.125 \text{ mol}) \times (8.314 \text{ J K}^{-1} \text{ mol}^{-1}) \times (273 \text{ K})}{(1.00 \times 1.013 \times 10^5 \text{ Nm}^{-2})}$$

$$= 0.0028 \text{ m}^3 = 2.8 \text{ L}$$

Note  $1 \text{ L} = 1 \text{ dm}^3 = 0.001 \text{ m}^3$ .

3. The pressure of the gas is the sum of the partial pressure of the sublimed carbon dioxide and the air.

The molar mass of  $\text{CO}_2$  is  $(12.01 + 2 \times 16.00) \text{ g mol}^{-1} = 44.01$ . Therefore, 7.8 g corresponds to:

$$\text{number of moles} = \frac{\text{mass}}{\text{molar mass}} = \frac{7.8 \text{ g}}{44.01 \text{ g mol}^{-1}} = 0.18 \text{ mol}$$

Using the perfect gas law,  $PV = nRT$  and assuming room temperature with  $T = 298 \text{ K}$ :

$$P_{\text{CO}_2} = \frac{nRT}{V} = \frac{(0.18 \text{ mol}) \times (8.314 \text{ J K}^{-1} \text{ mol}^{-1}) \times (298 \text{ K})}{(1.2 \times 10^{-3} \text{ m}^3)}$$

$$= 3.7 \times 10^5 \text{ Pa} = 3.6 \text{ atm}$$

As 760 mmHg is equivalent to 1 atm, the partial pressure of air is:

$$P_{\text{air}} = \frac{740}{760} = 0.974 \text{ atm}$$

The total pressure is the sum of the partial pressures so that:

$$P_{\text{total}} = P_{\text{air}} + P_{\text{CO}_2} = (0.974 + 3.6 \text{ atm}) = 4.6 \text{ atm}$$

4. The pressure of the gas is the sum of the partial pressure of nitrogen and water vapour.

As 760 mmHg is equivalent to 1 atm, the partial pressure of water vapour is:

$$P_{\text{water vapour}} = \frac{17.5}{760} = 0.0230 \text{ atm}$$

As the total pressure is 1 atm, the partial pressure of nitrogen is:

$$P_{\text{nitrogen}} = (1.00 - 0.0230) \text{ atm} = 0.98 \text{ atm}$$

Using the ideal gas law,  $PV = nRT$ , the number of moles of nitrogen corresponding to this pressure at  $20^\circ \text{C}$  is:

$$n = \frac{PV}{RT} = \frac{(0.98 \times 1.013 \times 10^5 \text{ Pa}) \times (645 \times 10^{-6} \text{ m}^3)}{(8.314 \text{ J K}^{-1} \text{ mol}^{-1}) \times ((20 + 273) \text{ K})} = 0.026 \text{ mol}$$

The molar mass of  $\text{N}_2$  is  $2 \times 14.01 = 28.02 \text{ g mol}^{-1}$ . Hence, this number of moles corresponds to a mass of:

$$\begin{aligned} \text{mass of nitrogen} &= \text{number of moles} \times \text{molar mass} \\ &= (0.026 \text{ mol}) \times (28.02 \text{ g mol}^{-1}) = 0.74 \text{ g} \end{aligned}$$