Answer on Question #56494-Physics-Other

1/ An oxygen cylinder of volume 30 liters has an initial gauge pressure of 15 atm and a temperature of 27°C. After some oxygen is withdrawn from the cylinder, the gauge pressure drops to 11 atm and its temperature drops to 17 °C. Estimate the mass of oxygen taken out of the cylinder.

Solution

($R = 8.31\, Jmol^{-1}K^{-1}$, molecular mass of $O_2 = 32\mu$).

Volume of oxygen, $V_1 = 30\ litres = 30 \cdot 10^{-3}\, m^3$.

Gauge pressure, $P_1 = 15\, atm = 15 \cdot 1.013 \cdot 10^5\, Pa$.

Temperature, $T_1 = 27°C = 300\, K$.

Universal gas constant, $R = 8.31\, Jmol^{-1}K^{-1}$.

Let the initial number of moles of oxygen gas in the cylinder be n_1.

The gas equation is given as:

$$P_1V_1 = n_1RT_1 \rightarrow n_1 = \frac{P_1V_1}{RT_1}$$

$$n_1 = \frac{(15.195 \cdot 10^5 \cdot 30 \cdot 10^{-3})}{(8.314 \cdot 300)} = 18.276.$$

But $n_1 = \frac{m_1}{M}$.

Where, m_1 is Initial mass of oxygen, $M = 32$ g is Molecular mass of oxygen. So,

$$m_1 = n_1M = 18.276 \cdot 32 = 584.84 \text{ g}$$

After some oxygen is withdrawn from the cylinder, the pressure and temperature reduces.

Volume, $V_2 = 30\ litres = 30 \cdot 10^{-3}\, m^3$.

Gauge pressure, $P_2 = 11\, atm = 11 \cdot 1.013 \cdot 10^5\, Pa$.

Temperature, $T_2 = 17°C = 290\, K$.

Let n_2 be the number of moles of oxygen left in the cylinder.

The gas equation is given as:

$$P_2V_2 = n_2RT_2 \rightarrow n_2 = \frac{P_2V_2}{RT_2}$$

$$n_2 = \frac{(11.143 \cdot 10^5 \cdot 30 \cdot 10^{-3})}{(8.314 \cdot 290)} = 13.86.$$

But $n_2 = \frac{m_2}{M}$.

Where, m_2 is the mass of oxygen remaining in the cylinder.

$$m_2 = n_2 M = 13.86 \cdot 32 = 453.1\ g.$$

The mass of oxygen taken out of the cylinder is given by the relation:

$$Initial\ mass\ of\ oxygen\ in\ the\ cylinder - Final\ mass\ of\ oxygen\ in\ the\ cylinder = m_1 - m_2$$
$$= 584.84\ \text{g} - 453.1\ \text{g} = 131.74\ \text{g} = 0.131\ \text{kg}.$$

Therefore, 0.131 kg of oxygen is taken out of the cylinder.

2/ A balloon whose volume is 750 m3 is to be filled with hydrogen at atmospheric pressure (1.01105 Pa).

(a) If the hydrogen is stored in cylinders with volumes of 1.90 m3 at a gauge pressure of 1.20 106 Pa, how many cylinders are required? Assume that the temperature of the hydrogen remains constant.

(b) What is the total weight (in addition to the weight of the gas) that can be supported by the balloon if the gas in the balloon and the surrounding air are both at 15.0°C? The density of air at 15.0oC and atmospheric pressure is 1.23 kg/m3.

(c) What weight could be supported if the balloon were filled with helium (molar mass 4.00 g/mol) instead of hydrogen at 15.0°C? What are your observations?

Solution

(a) The absolute pressure of the gas in a cylinder is:

$$P = 1.20 \cdot 10^6 + 1.01 \cdot 10^5 = 1.30 \cdot 10^6\ Pa.$$

At atmospheric pressure, the volume of hydrogen will increase by a factor of

$$\frac{1.30 \cdot 10^6}{1.01 \cdot 10^5}$$

so the number of cylinders is:

$$N = \frac{750\ m^3}{1.90\ m^3 \left(\frac{1.30 \cdot 10^6}{1.01 \cdot 10^5}\right)} = 31.$$

(b) The difference between the weight of the air displaced and the weight of hydrogen is:

$$W_{H_2} = (\rho_{air} - \rho_{H_2})Vg = \left(\rho_{air} - \frac{pM_{H_2}}{RT}\right)Vg$$
$$= \left(1.23\frac{kg}{m^3} - \frac{1.01 \cdot 10^5 Pa \cdot 2.2 \cdot 10^{-3}\frac{kg}{mol}}{8.314\frac{J}{molK} \cdot 288.15K}\right) 9.80\frac{m}{s^2} \cdot 750m^3 = 8.42 \cdot 10^3 N$$
$$= 8.42\ kN.$$

(c) The difference between the weight of the air displaced and the weight of helium is:

$$W_{He_2} = (\rho_{air} - \rho_{He_2})Vg = \left(\rho_{air} - \frac{pM_{He_2}}{RT}\right)Vg$$

$$= \left(1.23\frac{kg}{m^3} - \frac{1.01\cdot\ 10^5 Pa\cdot 4.00\cdot\ 10^{-3}\frac{kg}{mol}}{8.314\frac{J}{molK}\cdot 288.15K}\right)9.80\frac{m}{s^2}\cdot 750m^3 = 7.80\cdot 10^3 N$$

$$= 7.80\ kN.$$

It is less than for hydrogen.