What is the pressure in mmHg of a gas mixture that contains 1g of H_2 and 8g of Ar in 3.0L container at 27 degree Celsius?

Solution:

According to ideal gas law:

PV=nRT

(where P – total pressure, V – volume of container, n – total number of moles in container, R – universal gas constant, T – temperature in Kelvins).

Number of moles of Hydrogen:

$$n(H_2)=\frac{m(H_2)}{M(H_2)}=\frac{1\ g}{2\ g/mol}=0.5\ mol$$

Number of moles of Argon:

$$n(Ar)=\frac{m(Ar)}{M(Ar)}=\frac{8\ g}{40\ g/mol}=0.2\ mol$$

Let us find pressure in the container:

$$P=\frac{nRT}{V}=\frac{(0.5+0.2)mol*8.314\frac{L*kPa}{mol*K}*300K}{3L}=581.98\ kPa$$

At last, we will transform kPa into mmHg:

760 mmHg corresponds to 101.325 kPa;

x mmHg corresponds to 581.98 kPa;

$$x=\frac{581.98\ kPa*760\ mmHg}{101.325\ kPa}=4365.2\ \text{mmHg};$$

Answer:

Pressure in the container is equal to 4365.2 mmHg.