

Why Do Hot Air Balloons Float?

THE IDEAL GAS LAW

$$n = \frac{m}{M}$$

$$D = \frac{m}{V}$$

$PV = nRT$ where

P = pressure in atmosphere
 V = volume in liters
 n = number of moles of gas
 R = Universal Gas Constant = 0.0821 L·atm/mol·K
 T = Kelvin temperature

M = molar mass
 m = mass
 D = density

1. How many moles of oxygen will occupy a volume of 2.50 liters at 1.20 atm and 25 °C?

0.123 mol $n = \frac{PV}{RT} = \frac{(1.20 \text{ atm})(2.50 \text{ L})}{(0.0821 \text{ L}\cdot\text{atm/mol}\cdot\text{K})(298 \text{ K})} = 0.123 \text{ mol}$

2. What volume will 2.00 moles of nitrogen occupy at 720. mmHg and 20. °C? $P = \frac{720 \text{ mmHg}}{760 \text{ mmHg/atm}} = 0.947 \text{ atm}$

50.8 L $V = \frac{nRT}{P} = \frac{(2.00 \text{ mol})(0.0821 \text{ L}\cdot\text{atm/mol}\cdot\text{K})(293 \text{ K})}{(0.947 \text{ atm})} = 50.78 \text{ L}$

3. What pressure will be exerted by 25.0 g of CO₂ at temperature of 25 °C and a volume of 500. mL?

27.8 atm $P = \frac{mRT}{M \cdot V} = \frac{(25.0 \text{ g})(0.0821 \text{ L}\cdot\text{atm/mol}\cdot\text{K})(298 \text{ K})}{44.01 \text{ g/mol} (0.500 \text{ L})} = 27.796 \text{ atm}$

4. At what temperature will 5.00 g of Cl₂ exert a pressure of 900. mmHg at a volume of 750. mL?

153 K $T = \frac{PVM}{Rm} = \frac{(900. \text{ mmHg})(0.750 \text{ L})(70.91 \text{ g/mol})}{(62.4 \text{ mmHg}\cdot\text{L/mol}\cdot\text{K})(5.00 \text{ g})} = 153.41 \text{ K}$

5. What is the density of NH₃ at 800. mmHg and 25 °C? 0.733 g/L

$$D = \frac{PM}{RT} = \frac{(800 \text{ mmHg})(17.03 \text{ g/mol})}{(62.4 \text{ mmHg}\cdot\text{L/mol}\cdot\text{K})(298 \text{ K})} = 0.7327 \text{ g/L}$$

6. If the density of a gas is 1.2 g/L at 745 mmHg and 20. °C, what is its molar mass?

29 g/L $M = \frac{DRT}{P} = \frac{(1.2 \text{ g/L})(62.4 \text{ mmHg}\cdot\text{L/mol}\cdot\text{K})(293 \text{ K})}{745 \text{ mmHg}} = 29.45 \text{ g/L}$

7. How many moles of nitrogen gas will occupy a volume of 347 mL at 6680 mmHg and 27 °C?

0.124 mol N₂ $n = \frac{PV}{RT} = \frac{(6680 \text{ mmHg})(0.347 \text{ L})}{(62.4 \text{ mmHg}\cdot\text{L/mol}\cdot\text{K})(300 \text{ K})} = 0.1238 \text{ mol}$

8. What volume will 454 grams (1 lb) of hydrogen occupy at 1.05 atm and 25 °C?

5240 L $V = \frac{mRT}{M \cdot P} = \frac{(454 \text{ g})(0.0821 \text{ L}\cdot\text{atm/mol}\cdot\text{K})(298 \text{ K})}{(2.02 \text{ g/mol})(1.05 \text{ atm})} = 5236.9 \text{ L}$

9. Find the number of grams of CO₂ that exert a pressure of 785 mmHg at a volume of 32.5 L and a temperature of 32 °C. 59.0 g

$$m = \frac{PVM}{RT} = \frac{(785 \text{ mmHg})(32.5 \text{ L})(44.01 \text{ g/mol})}{(62.4 \text{ mmHg}\cdot\text{L/mol}\cdot\text{K})(305 \text{ K})} = 58.996 \text{ g}$$

10. An elemental gas has a mass of 10.3 g. If the volume is 58.4 L and the pressure is 758 mmHg at a temperature of 2.5 °C, what is the gas? Helium

$$M = \frac{mRT}{PV} = \frac{(10.3 \text{ g})(62.4 \text{ mmHg}\cdot\text{L/mol}\cdot\text{K})(275.5 \text{ K})}{(758 \text{ mmHg})(58.4 \text{ L})}$$

$M = 4.00$ So gas is Helium