

Chemistry 1000 – In Class Gas Law Problems - SOLUTIONS

1. What volume does 0.64 mol of oxygen occupy at STP?

$$0.64 \text{ mol} \times 22.41 \text{ L/mol} = 14 \text{ L}$$

2. Calculate the temperature of 1.82 mol of a gas contained in a vessel of 5.43 L on which it exerts a pressure of 9.42 atm.

$$p = 9.42 \text{ atm} \times 101325 \text{ Pa/atm} = 954000 \text{ Pa}$$

$$V = 5.43 \text{ L} \times 0.001 \text{ m}^3/\text{L} = 0.00543 \text{ m}^3$$

$$\begin{aligned} T &= pV/nR = 954000 \text{ Pa} \times 0.00543 \text{ m}^3 / (1.82 \text{ mol} \times 8.3145 \text{ J K}^{-1} \text{ mol}^{-1}) \\ &= 354 \text{ Pa m}^3 \text{ K J}^{-1} = 354 (\text{kg m}^{-1} \text{ s}^{-2}) \text{ m}^3 \text{ K} (\text{kg}^{-1} \text{ m}^2 \text{ s}^{-2}) \\ &= 354 \text{ K} = 81 \text{ }^\circ\text{C} \end{aligned}$$

3. A gas has a volume of 0.78 L when measured at 20.1 °C and 1.00 atm. What is the volume of the gas at 36.5 °C at 1.00 atm pressure?

p and $n = \text{const.}$

$$V_1/T_1 = V_2/T_2$$

$$V_2 = V_1/T_1 \times T_2 = (0.78 \text{ L}/20.1 \text{ }^\circ\text{C}) \times 36.5 \text{ }^\circ\text{C} = 1.4 \text{ L}$$

4. The density of air at 0 °C and a pressure of 101325 Pa is 0.001293 g cm⁻³. What is the apparent molar mass of air, assuming it behaves like an ideal gas.

$$T = 273 \text{ K}$$

$$\rho = 0.001293 \text{ g cm}^{-3} \times 1000000 \text{ cm}^3 / 1 \text{ m}^3 = 1293 \text{ g m}^{-3} \text{ [we are not converting g into the SI unit kg, since we need the unit g/mol for the molar mass]}$$

$$\begin{aligned} M &= \rho RT/p = 1293 \text{ g m}^{-3} \times 8.3145 \text{ J K}^{-1} \text{ mol}^{-1} \times 273 \text{ K} / 101325 \text{ Pa} \\ &= 29.0 \text{ g m}^{-3} \text{ J mol}^{-1} \text{ Pa}^{-1} = 29.0 \text{ g m}^{-3} (\text{kg}^1 \text{ m}^2 \text{ s}^{-2}) \text{ mol}^{-1} (\text{kg}^{-1} \text{ m s}^2) \\ &= 29.0 \text{ g/mol} \end{aligned}$$

5. A 2.1 L vessel contains 4.65 g of a gas at 1.00 atm and 27.0 °C. Calculate the density and the molar mass of the gas.

$$T = (27.0 + 273.15) \text{ K} = 300.2 \text{ K}$$

$$p = 1 \text{ atm} \times 101325 \text{ Pa/atm} = 101325 \text{ Pa}$$

$$V = 2.1 \text{ L} = 2100 \text{ cm}^3 = 0.0021 \text{ m}^3$$

$$\rho = 4.65 \text{ g}/2100 \text{ cm}^3 = 0.00221 \text{ g cm}^{-3} = 2210 \text{ g m}^{-3}$$

$$\begin{aligned} M &= \rho RT/p = 2210 \text{ g m}^{-3} \times 8.3145 \text{ J K}^{-1} \text{ mol}^{-1} \times 300.2 \text{ K} / 101325 \text{ Pa} \\ &= 54.4 \text{ g m}^{-3} \text{ J mol}^{-1} \text{ Pa}^{-1} = 54.4 \text{ g m}^{-3} (\text{kg}^1 \text{ m}^2 \text{ s}^{-2}) \text{ mol}^{-1} (\text{kg}^{-1} \text{ m s}^2) \\ &= 54.4 \text{ g/mol} \end{aligned}$$

6. An anesthetic contains 64.9 %C, 13.5 %H, and 21.6%O. One liter of the gas at 120 °C and 750 mmHg weighs 2.30 g. What is the molecular formula of the gas?

$$64.9 \text{ g C} - 5.40 \text{ mol C}$$

$$13.5 \text{ g H} - 13.4 \text{ mol H}$$

$$21.6 \text{ g O} - 1.35 \text{ mol O}$$

$$\text{C:H:O} \approx 4:10:1$$

$$\text{Empirical formula: } \text{C}_4\text{H}_{10}\text{O}$$

$$M(C_4H_{10}O) = 74.123 \text{ g/mol}$$

$$T = (120 + 273)K = 393 \text{ K}$$

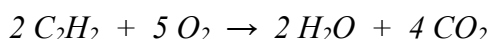
$$p = 750 \text{ Torr} \times 101325 \text{ Pa} / 760 \text{ Torr} = 100000 \text{ Pa}$$

$$V = 1L = 0.001 \text{ m}^3$$

$$\begin{aligned} M &= m RT/pV = 2.30 \text{ g} \times 8.3145 \text{ J K}^{-1} \text{ mol}^{-1} \times 393 \text{ K} / (100000 \text{ Pa} \times 0.001 \text{ m}^3) \\ &= 75.2 \text{ g J mol}^{-1} \text{ Pa}^{-1} \text{ m}^{-3} = 75.2 \text{ g (kg}^1 \text{ m}^2 \text{ s}^{-2}) \text{ mol}^{-1} (\text{kg}^{-1} \text{ m s}^2) \text{ m}^{-3} \\ &= 75.2 \text{ g/mol} \end{aligned}$$

molecular formula: $C_4H_{10}O$

7. Calculate the volume of $O_{2(g)}$ (in L) at STP required for the complete combustion of 2.64 L of acetylene ($C_2H_{2(g)}$) at STP.

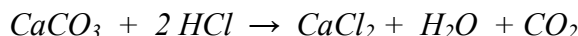


$$\text{mol } O_2 = \text{mol } C_2H_2 \times 5 \text{ mol } O_2 / 2 \text{ mol } C_2H_2$$

Since the number of moles and the volume are proportional when p and $T = \text{const.}$, we can substitute mol by volume in the above equation.

$$V(O_2) = L O_2 = L C_2H_2 \times 5 L O_2 / 2 L C_2H_2 = 2.64 L C_2H_2 \times 5 L O_2 / 2 L C_2H_2 = 6.60 L O_2$$

8. Calculate the volume of $CO_{2(g)}$ produced at STP that could be obtained by allowing 45.0 g of $CaCO_{3(s)}$ to react with hydrochloric acid.



$$M(CaCO_3) = 100.087 \text{ g mol}^{-1}$$

$$n(CaCO_3) = 45.0 \text{ g} / (100.087 \text{ g mol}^{-1}) = 0.450 \text{ mol}$$

$$\begin{aligned} n(CO_2) &= 0.450 \text{ mol } CaCO_3 \text{ consumed} \times 1 \text{ mol } CO_2 \text{ formed} / 1 \text{ mol } CaCO_3 \text{ consumed} \\ &= 0.450 \text{ mol } CO_2 \text{ formed} \end{aligned}$$

At STP 1 mol: 22.41 L

$$\text{At STP: } V(CO_2) = 0.450 \text{ mol} \times 22.41 \text{ L/mol} = 10.1 \text{ L}$$

1 mol of an ideal gas occupies a volume of 22.41 L (0.02241 m^3) at 0°C (273.15 K) and 1 atm of pressure (101325 Pa).

$$pV = nRT$$

$$n = m/MW$$

$$\rho = m/V$$

$$R = 8.3145 \text{ J K}^{-1} \text{ mol}^{-1}$$