

Please note that both mm Hg and atm are used as units of pressure. It is a simple conversion: $760 \text{ mm} = 1 \text{ atm}$

- Given 3.43 g of gas in a 2.00 L container at 25.0°C and a pressure of 1140 mm Hg:
 - Determine the number of moles of gas in the container.
 - Recalling that molar mass (molecular weight) is nothing more than a quotient of grams per mole (mass/moles), determine the molar mass of this gas.
 - What might be the identity of this gas?
- A 3.0 L flask at 30.0°C contains 0.250 mole of Cl₂ gas.
 - What is the pressure in the flask?
 - What is the mass of the gas in the flask?
 - What is the density of the chlorine gas in this flask?
- A 500.0 ml flask contained O₂ gas at 25.0°C at a pressure of 4.5 atm.
 - What is the number of moles in the flask?
 - What is the mass of the gas in the flask?
 - What is the density of the oxygen in the flask?
- A 5.0 L flask of carbon dioxide gas at a pressure of 4.54 atm had a mass of 36 g?
 - How many moles of gas are in this flask?
 - What is the temperature, in Kelvin and °C, of the gas in this flask?
- How large of a metal gas canister would you need to contain 20.0 moles of compressed gas at a pressure of 22 atm and at room temperature, 25.0°C?
- The density of SO₂ gas in a container at room temperature, 25.0°C is 2.51 g/L.
 - Determine the pressure in this flask.
- Determine the density of O₂ at STP.
- A 5.0 L flask at 60.0°C contains 0.055 mole of oxygen gas.
 - What is the pressure in the flask?
 - What is the mass of the gas in the flask?
 - What is the density of the oxygen gas in this flask?
- Determine the molar mass of gas in a container at -50.0°C and 6 atm pressure with a density of 14.5 g/L.

1. Again apply the ideal gas law solving for n. Be sure your temp is in Kelvin and select the R that matches the P units. Remember that molar mass is mass/moles.

$$a. \quad n = \frac{PV}{RT} \quad n = \frac{(1140 \text{ mmHg})(2L)}{(62.4 \frac{\text{mmHg} \cdot L}{\text{mol} \cdot K})(298K)} \quad n = 0.123 \text{ mole}$$

$$b. \quad \text{Since } MM = \frac{\text{mass}}{\text{moles}} \quad \text{thus } MM = \frac{3.43\text{g}}{0.123\text{mol}} \quad MM = 27.9 \text{ g/mole}$$

c. It's likely to be diatomic nitrogen, N_2 with $MM = 28 \text{ g/mole}$

2. Apply the ideal gas law: $PV = nRT$ solving for P. Depending which R value you use, tells you which label to put on your resulting pressure value. Don't forget that your temp must be in Kelvin.

$$a. \quad P = \frac{(0.25\text{mol})(62.4 \frac{\text{mmHg} \cdot L}{\text{mol} \cdot K})(303K)}{(3L)} \quad \text{thus } P = 1577 \text{ mmHg then round off to } 1600 \text{ mmHg}$$

$$\text{OR } P = \frac{(0.25\text{mol})(0.821 \frac{\text{atm} \cdot L}{\text{mol} \cdot K})(303K)}{(3L)} \quad \text{thus } P = 2.1 \text{ atm}$$

$$b. \quad \text{Since } 0.25\text{mol} \left(\frac{71\text{g}}{1\text{mol}} \right) = 17.8 \text{ g}$$

$$c. \quad \text{Remember that } D = \frac{\text{mass}}{\text{vol}} \quad \text{so } D = \frac{17.8\text{g}}{3L} = 5.9 \text{ g/L}$$

3. Apply the ideal gas law: $PV = nRT$ solving for n. Here you can use the pressure given in atm, but you must choose the 0.0821 atm L/mole K gas constant. Or you can change 4.5 atm to 3420 mm Hg and use the 62.4 mmHg L/mole K gas constant. You must change 500 ml to 0.5 L because both gas constants have units in L, and so the volume must be in liters so it can cancel out. And of course, you must change the temperature to Kelvin.

$$a. \quad n = \frac{(4.5\text{atm})(0.5L)}{(0.0821 \frac{\text{atm} \cdot L}{\text{mol} \cdot K})(298K)} = 0.092 \text{ mole}$$

$$b. \quad 0.092\text{mol} \left(\frac{32\text{g}}{1\text{mol}} \right) = 2.9 \text{ g}$$

$$c. \quad \text{Remember that } D = \frac{\text{mass}}{\text{vol}} \quad \text{so } D = \frac{2.9\text{g}}{0.5L} = 5.9 \text{ g/L}$$

4. First change mass to moles using the molar mass of CO_2 the apply the ideal gas law, $PV = nRT$ solving for T. Remember that the answer will come out in Kelvin, and you must change to report it in Celsius.

$$a. \quad 366\text{g} \left(\frac{1\text{mol}}{44\text{g}} \right) = 0.82 \text{ mole}$$

$$b. \quad T = \frac{PV}{nR} \quad T = \frac{(4.54\text{atm})(5L)}{(0.82\text{mol})(0.0821 \frac{\text{atm} \cdot L}{\text{mol} \cdot K})} = 337 \text{ K which converts to } 64^\circ\text{C}$$

5. Use the ideal gas law, $PV = nRT$ and solve for V. Be sure you convert your temp to Kelvin.

$$V = \frac{(20\text{mol})(0.821 \frac{\text{atm} \cdot L}{\text{mol} \cdot K})(298K)}{(22\text{atm})} = 22.2 \text{ L}$$

6. Use the molar mass Kitty Cat !! $MM = \frac{dRT}{P}$ and solve for P

$$a. \quad P = \frac{(2.51 \frac{g}{L})(62.4 \frac{mmHg \cdot L}{mol \cdot K})(298K)}{(64 \frac{g}{mol})} = 732 \text{ mm Hg which is also equal to } 0.9663 \text{ atm}$$

7. Take the simple route by dividing the molar mass, 32 g/mole by the molar volume 22.4 L/mole to get 1.42 g/L

8. Use $PV = nRT$ and solve for P

$$a. \quad P = \frac{(0.055 \text{ mol})(0.821 \frac{atm \cdot L}{mol \cdot K})(333K)}{(5L)} = 0.30 \text{ atm or } 228 \text{ mmHg}$$

$$b. \quad 0.055 \text{ mol} \left(\frac{32g}{1 \text{ mol}} \right) = 1.76 \text{ g} \quad \text{since we know it is oxygen}$$

$$c. \quad \text{Since } D = \frac{\text{mass}}{\text{vol}} \quad \text{thus } D = \frac{1.76g}{5L} = 0.352 \text{ g/L}$$

$$9. \quad MM = \frac{dRT}{P} \quad MM = \frac{(14.5 \frac{g}{L})(0.0821 \frac{atm \cdot L}{mol \cdot K})(223K)}{(6atm)} = 44g/mol$$