P D.2 (pg 1 of 3) **Ideal Gas Law & Density**

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

1. Given 3.43 g of gas in a 2.00 L container at 25.0ºC and a pressure of 1140 mm Hg:
   a. Determine the number of moles of gas in the container.
   b. 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.
   c. What might be the identity of this gas?

2. A 3.0 L flask at 30.0ºC contains 0.250 mole of Cl_2 gas.
   a. What is the pressure in the flask?
   b. What is the mass of the gas in the flask?
   c. What is the density of the chlorine gas in this flask?

3. A 500.0 ml flask contained O_2 gas at 25.0ºC at a pressure of 4.5 atm.
   a. What is the number of moles in the flask?
   b. What is the mass of the gas in the flask?
   c. What is the density of the oxygen in the flask

4. A 5.0 L flask of carbon dioxide gas at a pressure of 4.54 atm had a mass of 36 g?
   a. How many moles of gas are in this flask?
   b. What is the temperature, in Kelvin and ºC, of the gas in this flask?

5. 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, 25.0ºC?

6. The density of SO_2 gas in a container at room temperature, 25.0ºC is 2.51 g/L.
   a. Determine the pressure in this flask.

7. Determine the density of O_2 at STP.

8. A 5.0 L flask at 60.0ºC contains 0.055 mole of oxygen gas.
   a. What is the pressure in the flask?
   b. What is the mass of the gas in the flask?
   c. What is the density of the oxygen gas in this flask?

9. Determine the molar mass of gas in a container at −50.0ºC 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. $n = \frac{PV}{RT}$ $\quad n = \frac{(1140 mmHg)(2L)}{(62.4 \frac{mmHg \bullet L}{mol \bullet K})(298K)}$ $\quad$ n = 0.123 mole

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

   c. It's likely to be diatomic nitrogen, N_2 with MM = 28 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. $P = \frac{(0.25mol)(62.4 \frac{mmHg \bullet L}{mol \bullet K})(303K)}{(3L)}$ $\quad$ thus P = 1577 mmHg then round off to 1600 mmHg

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

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

   c. Remember that $D = \frac{mass}{vol}$ so $D = \frac{17.8g}{3L}$ = 5.9 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. $n = \frac{(4.5atm)(0.5L)}{(0.0821 \frac{atm \bullet L}{mol \bullet K})(298K)}$ = 0.092 mole

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

   c. Remember that $D = \frac{mass}{vol}$ so $D = \frac{2.9g}{0.5L}$ = 5.9 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. $366g\left(\frac{1mol}{44g}\right)$ = 0.82 mole

   b. $T = \frac{PV}{nR}$ $\quad T = \frac{(4.54atm)(5L)}{(0.82mol)(0.0821 \frac{atm \bullet L}{mol \bullet K})}$ = 337 K which converts to 64ºC

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

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

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

   a. $P = \dfrac{(2.51\frac{g}{L})(62.4\frac{mmHg \bullet L}{mol \bullet K})(298K)}{(64\frac{g}{mol})}$ = 732 mm Hg which is also equal to 0.9663 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. $P = \dfrac{(0.055mol)(0.821\frac{atm \bullet L}{mol \bullet K})(333K)}{(5L)}$ = 0.30 atm or 228 mmHg

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

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

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