

Gasses:

Units of pressure

torr = Torr

atm = atmosphere

mmHg = millimeters of mercury

bar = bar

kPa = kilopascal

lb/in² = pounds per square inch = psi

convert to atm

$$760 \text{ torr} = 1 \text{ atm}$$

$$1 \text{ atm} = 1 \text{ atm}$$

$$760 \text{ mmHg} = 1 \text{ atm}$$

$$1.01325 \text{ bar} = 1 \text{ atm}$$

$$101.325 \text{ kPa} = 1 \text{ atm}$$

$$14.7 \text{ lb/in}^2 = 1 \text{ atm}$$

A geochemist heats a limestone sample (CaCO₃) and collects the CO₂ released. After the system comes to room temp, $\Delta h = 291.4 \text{ mmHg}$. Calculate P_{CO₂} in torr, atm, and kPa.

$$P_{\text{CO}_2} (\text{torr}) = 291.4 \text{ mmHg} \times \frac{1 \text{ torr}}{1 \text{ mmHg}} = 291.4 \text{ torr}$$

$$P_{\text{CO}_2} (\text{kPa}) = 0.3834 \text{ atm} \times \frac{101.325 \text{ kPa}}{1 \text{ atm}}$$

$$P_{\text{CO}_2} (\text{atm}) = 291.4 \text{ torr} \times \frac{1 \text{ atm}}{760 \text{ torr}} = 0.3834 \text{ atm}$$

$$= 38.95 \text{ kPa}$$

Gas laws:

Boyle's Law - vary P, fixed n and T

Charles' Law - vary T, fixed n and P

Avogadro's Law - vary n, fixed P and T

Ideal gas law $PV = nRT$

P = pressure (atm) n = moles (mol) T = temp (K)

V = volume (L) R = gas constant ($\frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}$)

When solving gas law problems, keep in mind two types

1. A change in one of the four variables causes a change in another, while the two remaining variables remain constant.

use: Gas law reduced to one of the individual laws.

2. One variable is unknown, but the other three are known and no change occurs.

use: Ideal gas law

Examples:

A student finds the air trapped in a J tube to occupy 24.8 cm³ at 1.12 atm. Mercury is added to the tube causing the pressure to increase to 2.64 atm. What is the new volume of air in (L)?

Relationship: Volume - pressure

Looks like: Boyle

Solve:

$$P_1 = 1.12 \text{ atm}$$

$$P_2 = 2.64 \text{ atm}$$

$$\frac{P_1 V_1}{P_1 T_1} = \frac{P_2 V_2}{P_2 T_2} \Rightarrow P_1 V_1 = P_2 V_2$$

$$V_1 = 24.8 \text{ cm}^3 = 0.0248 \text{ L}$$

$$V_2 = ?$$

$$(1.12 \text{ atm})(0.0248 \text{ L}) = (2.64 \text{ atm})(V_2)$$

$$V_2 = 0.0105 \text{ L}$$

Steel fuel delivery tanks have safety valves that open if internal pressure exceeds $1.00 \times 10^3 \text{ torr}$. When filled with methane at 23 deg C at 0.991 atm and placed in a 100.00 deg C boiling water bath, does the safety valve open?

Relationship: Pressure - Temperature

Looks like: analogous to Charles (Fixed P & n) $\frac{P_1}{T_1} = \frac{P_2}{T_2} \Rightarrow \frac{753 \text{ torr}}{296 \text{ K}} = \frac{P_2}{373 \text{ K}}$
 Solve:
 $P_1 = 0.991 \text{ atm} = 753 \text{ torr}$ $P_2 = ?$
 $T_1 = 23^\circ\text{C} = 296 \text{ K}$ $T_2 = 100^\circ\text{C} = 373 \text{ K}$
 $P_2 = 949 \text{ torr}$
 No, the valve will not open

A model builder knows it takes 55.0 dm^3 of He to cause a scale model of a blimp to rise. His son fills his dad's blimp with 1.10 mol He , 26.2 dm^3 , and it does not rise. The son doesn't want to waste his dad's helium. How many more grams of He does he need to cause the blimp to rise?

Relationship: Volume - amount
 Looks like: Avogadro
 Solve:
 $n_1 = 1.10 \text{ mol}$ $n_2 = ?$
 $V_1 = 26.2 \text{ dm}^3$ $V_2 = 55.0 \text{ dm}^3$
 $\frac{V_1}{n_1} = \frac{V_2}{n_2} \Rightarrow \frac{1.10 \text{ mol}}{26.2 \text{ dm}^3} = \frac{n_2}{55.0 \text{ dm}^3}$ $n_2 = 2.31 \text{ mol He}$
 $2.31 \text{ mol He} - 1.10 \text{ mol He} = 1.21 \text{ mol He needed}$
 $1.21 \text{ mol He} \times \frac{4.003 \text{ g He}}{1 \text{ mol He}} = 4.84 \text{ g He}$

An O_2 storage tank has a capacity of 438 L . It is filled with 0.885 kg O_2 (MW = 32.00 g/mol). At 21 deg C how much pressure does the O_2 exert?

Relationship: Unknown Variable @ fixed conditions
 Looks like: Ideal gas law
 Solve:
 $V = 438 \text{ L}$ $R = 0.0821 \frac{\text{L}\cdot\text{atm}}{\text{K}\cdot\text{mol}}$ $T = 21^\circ\text{C} = 294 \text{ K}$ $P(438 \text{ L}) = (27.7 \text{ mol})(0.0821)(294 \text{ K})$
 $n = 0.885 \text{ kg O}_2 = 27.7 \text{ mol O}_2$ $P = ?$ $P = 1.53 \text{ atm}$

A student doing research using the tank in the above question does not close the valve completely when he is finished with his experiment. By the time someone else comes along, the tank reads 1.37 atm . How much O_2 was lost by the careless student?

$(1.37 \text{ atm})(438 \text{ L}) = n(0.0821 \frac{\text{L}\cdot\text{atm}}{\text{K}\cdot\text{mol}})(294 \text{ K})$ $0.885 \text{ kg O}_2 - 0.795 \text{ kg O}_2 = 0.089474 \text{ kg O}_2 \text{ lost}$
 $n = 24.86 \text{ mol O}_2$
 $24.86 \text{ mol O}_2 \times \frac{32.00 \text{ g}}{\text{mol}} = 795.52 \text{ g O}_2 = 0.795526 \text{ kg O}_2$

Density of a gas can be found using the ideal gas law by:

$d = \frac{m}{V}$ $n = \frac{m}{\text{MW}}$ $PV = nRT \Rightarrow PV = \left(\frac{m}{\text{MW}}\right)RT \Rightarrow \frac{m}{V} = d = \frac{(\text{MW})(P)}{RT}$

To apply a green chem approach, a chemical engineer uses waste CO_2 from a manufacturing process instead of CFC's while making polystyrene containers. Find the density (g/L) of CO_2 and number of molecules per liter at STP (0 deg C and 1 atm) and at room temp (20.0 deg C and 1.00 atm).

STP $T = 0^\circ\text{C} = 273 \text{ K}$ $P = 1 \text{ atm}$ $\text{MW CO}_2 = 44.01 \text{ g/mol}$
 $d = \frac{(44.01 \text{ g/mol})(1.00 \text{ atm})}{(0.0821 \frac{\text{L}\cdot\text{atm}}{\text{K}\cdot\text{mol}})(273 \text{ K})} = 1.96 \text{ g/L}$

$\frac{\text{molecules CO}_2}{\text{L}} = \frac{1.96 \text{ g CO}_2}{\text{L}} \times \frac{1 \text{ mol CO}_2}{44.01 \text{ g}} \times \frac{6.022 \times 10^{23} \text{ molecules}}{\text{mol}} = 2.68 \times 10^{22}$

Room temp $T = 20.0^\circ\text{C} = 293 \text{ K}$ $P = 1.00 \text{ atm}$ $\text{MW} = 44.01 \text{ g/mol}$

$d = 1.83 \text{ g/L}$

Calc same as STP, but w/ 293 K instead of 273 K .

$\frac{\text{molecules CO}_2}{\text{L}} = 2.50 \times 10^{22}$

$m = \text{mass}$
 $\text{MW} = \text{molecular weight or molar mass}$

- The density of a gas is directly proportional to its molar mass (Avogadro's Law...)

- The density of a gas is inversely proportional to the temperature (Charles' Law...)

Molar mass can be found by:

$$n = \frac{m}{MW} = \frac{PV}{RT} \Rightarrow MW = \frac{mRT}{PV} \text{ or } MW = \frac{dRT}{P}$$

An organic chemist isolates a colorless liquid from a petroleum sample. She obtains the following data:

V (flask) = 213 mL, T = 100.0 deg C, P = 754 torr, mass of flask + gas = 78.416 g, mass of flask = 77.834 g.

Calculate the molar mass of the liquid.

$$m = 78.416 \text{ g} - 77.834 \text{ g} = 0.582 \text{ g} = \text{mass of gas}$$

$$P_{(\text{atm})} = 754 \text{ torr} \times \frac{1 \text{ atm}}{760 \text{ torr}} = 0.992 \text{ atm}$$

$$V = 213 \text{ mL} = 0.213 \text{ L}$$

$$T = 100.0^\circ \text{C} = 373.2 \text{ K}$$

$$MW = \frac{mRT}{PV} = \frac{(0.582 \text{ g}) \left(0.0821 \frac{\text{L} \cdot \text{atm}}{\text{K} \cdot \text{mol}}\right) (373.2 \text{ K})}{(0.992 \text{ atm})(0.213 \text{ L})}$$

$$= 84.4 \text{ g/mol}$$