Pressure

$$P = \frac{F}{A}$$

$$P = \frac{kg \dfrac{m}{s^2}}{m^2}$$

$$P = \frac{m \bullet a}{A}$$

$$P = \frac{kg}{m \ s^2}$$

Pressure is force divided by area. The force we are talking about is created by the attraction of a column of air to the earth by gravity.

$$P = \frac{m \bullet g}{A}$$

Since volume is area times height, or rearranging A = V/h, we can substitute for A, area.

$$P = \frac{m \bullet g \bullet h}{V}$$

$$P = \frac{m}{V} \ g \bullet h$$

Mass divided by volume is just density.

$$P = d \bullet g \bullet h$$

This is how the only SI unit for pressure is determined. Torr, mmHg, and atm are all arbitrary units, but the pascal is based on SI units.

For example:

The atmosphere can support a column of mercury that measures 760 mm tall.

$$P = 13.6 \text{ g/cm}^3 \bullet 9.8 \text{ m/s}^2 \bullet 760 \text{ mmHg}$$

$$P = \left(\frac{13.6\ \cancel{g}}{1\ \cancel{cm^3}}\right)\left(\frac{1\ kg}{1000\ \cancel{g}}\right)\left(\frac{1 \times 10^6\ \cancel{cm^3}}{1\ \cancel{m^3}}\right)\left(\frac{9.8\ \cancel{m}}{s^2}\right)(0.760\ \cancel{m})$$

$$P = 101292.8\ \frac{kg}{m\ s^2}$$

$$P = 101292.8\ \text{Pa}$$

So,

$$1 \text{ atm} = 760 \text{ torr} = 101929.8 \text{ Pa}$$

Gases

The ideal gas Law

$$PV = nRT$$

Where does this come from?

Robert Boyle found

$$PV = \text{a constant}$$

That is, the product of the pressure of a gas times the volume of a gas is a constant for a given sample of gas. In Boyle's experiments the temperature (T) did not change, nor did the number of moles (n) of gas present. So, Boyle found

$$PV = (nRT)$$

but did not explore the effect the temperature, or the number of moles would have on pressure and volume.

Jaques Charles found

$$V = (\text{a constant}) \, T$$

That is, the volume of a given sample of gas increases linearly with the temperature if the pressure (P) and the amount of the gas (n) is kept constant. So Charles found

$$V = (nR/P) \, T$$

Avagadro's Postulate

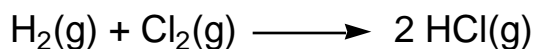At the same temperature and pressure equal volumes of all

gases contain the same number of molecules.

$$V = n \text{ (a constant)}$$

Or
$$V = n \, (RT/P)$$

Guy Lussac found that 1 volume of Cl_2 combined with 1 volume of H_2 to make 2 volumes of HCl. The equation for the reaction is

$$H_2(g) + Cl_2(g) \longrightarrow 2\,HCl(g)$$

With this example we can clearly see the relationship between the number of moles of a gas and the volume of a gas.

Not so coincidentally if V is held constant instead of P then

$$P = n \, (RT/V)$$

A **PV = nRT problem**

What is the volume of 1 mole of an ideal gas at STP (Standard Temperature and Pressure = 0 ˚C, 1 atm)?

$$PV = nRT$$

$$(1) \, V = 1(0.08206)(273.15)$$

$$V = 22.41 \, L$$

So, the volume of an ideal gas is 22.41 L/mol at STP.

This is probably one of the most remembered, and least useful numbers in chemistry.

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What is the volume of 5.0 g NH_3 at 25 ˚C and 1.0 atm pressure ?

can you use the volume of an ideal gas conversion that we determined above?

$$5.0 \, g \, NH_3 \quad \times \quad \frac{1 \, mol \, NH_3}{17.03 \, g \, NH_3} \quad \times \quad \frac{22.41 \, L \, NH_3}{1 \, mol \, NH_3}$$ **NO**

No, you cannot use the 22.4 L/mol because this number is only true for gases at 0 ˚C and 1 atm!

So, instead...
Find the number of moles of NH_3 present

$$5.0 \, \cancel{g \, NH_3} \quad \times \quad \frac{1 \, mol \, NH_3}{17.03 \, \cancel{g \, NH_3}} \quad = \quad 0.2936 \, mol \, NH_3$$

and use $PV=nRT$

$$(1 \, atm) \, V = (0.2936 \, mol) \, (0.08206 \, L•atm•mol^{-1}•K^{-1}) \, (298.15 \, K)$$

$$V = 7.18 = 7.2 \, L$$

Seltzer water is made by dissolving CO_2 in water. Seltzer can be
made at home using small containers of pressurized CO_2. If one
of the cartridges contains 20.00 mL CO_2 at 55.00 atm at 23.0 ˚C
and it expands into an empty seltzer bottle with a volume of
1.000 L and the resulting pressure is 1.000 atm what is the
temperature of the gas after it expands?

Before the gas expands...

$$(55 \text{ atm})(0.020 \text{ L}) = n \text{ R } (296.15 \text{ K})$$

and

$$\frac{(55 \text{ atm})(0.020 \text{ L})}{(296.15 \text{ K})} = n \text{ R}$$

after the gas expands...

$$(1.0 \text{ atm})(1.020 \text{ L}) = n \text{ R T}$$

or

$$\frac{(1.0 \text{ atm})(1.020 \text{ L})}{T} = n \text{ R}$$

since R has not changed (it is called the universal gas CONSTANT for a reason) and we have
not changed the number of moles of CO_2...

$$\frac{(1.0 \text{ atm})(1.020 \text{ L})}{T} = n \text{ R} = \frac{(55 \text{ atm})(0.020 \text{ L})}{(296.15 \text{ K})}$$

$$\frac{(1.0 \text{ atm})(1.020 \text{ L})}{T} = \frac{(55 \text{ atm})(0.020 \text{ L})}{(296.15 \text{ K})}$$

$$T = 274.61 \text{ K}$$

or...

$$T = 274.61 - 273.15 = 1.5 \text{ ˚C}$$

What will the pressure be when the gas warms to 23.0 ˚C?

before expansion and warming...
$$(55 \text{ atm})(0.020 \text{ L}) = n \, R \, (296.15 \text{ K})$$

after expansion and warming
$$(P)(1.020 \text{ L}) = n \, R \, (296.15 \text{ K})$$

once again R and n have not changed, and now not even T has not changed so...

$$\frac{(P)(1.020 \text{ L})}{(296.15 \text{ K})} = \frac{(55 \text{ atm})(0.020 \text{ L})}{(296.15 \text{ K})}$$

$$P = 1.078 \text{ atm}$$

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A simple experiment can be performed to determine the molar mass of an unknown volatile liquid. 1.5021 g of a volatile liquid was placed in a container which had a small pin-hole in the top, and the container was placed in boiling water. After all of the liquid had vaporized and the vapor stopped escaping from the pin-hole the container was cooled back to room temperature. 0.7523 g of the volatile liquid remained in the container, the volume of the container is 0.500 L, and the atmospheric pressure on that day was 756 torr. Determine the molar mass of this compound.

To determine the molar mass the number of moles in a given mass needs to be determined.

So, the question we have to answer is how many moles of liquid are present in a sample of liquid.

The key here is that at one point in the experiment all of the liquid was converted to gas.

"After all of the liquid had vaporized the container was cooled back to room temperature. 0.7523 g of the volatile liquid remained in the container"

So, we have alot of information about the unknown.

We know
• The mass of the unknown gas in the container is 0.7523 g.
   After all the liquid was converted to vapor the container was cooled and the vapor which was in the container condensed to a liquid. The mass of the liquid can be determined by weighing it on a balance.

• The volume that the gas occupies is 0.500 L
   As the liquid vaporizes it pushes the air out of the container. When all the liquid has vaporized only the unknown vapor is in the container. The volume of the container is 0.500 L.

• The pressure of the gas is 756 torr or 0.9947 atm.
   Since there is a pinhole in the container the gas on the inside of the container is at the same pressure as the gas outside the container. The gas outside the container is the atmosphere, and the pressure of the atmosphere was 756 torr.

• The temperature of the gas is 100 °C, or 373.15 K.
   While the container was place in boiling water the vaporized unknown filled the container. So, the temperature of the gas was the same as the temperature of the boiling water.

So,
$$PV=nRT$$

We know P, V, and T! We can find n.

$$(0.9947 \text{ atm})(0.500 \text{ L}) = n(0.08206 \text{ L} \cdot \text{atm} \cdot \text{mol}^{-1} \cdot \text{K}^{-1})(373.15 \text{ K})$$

n = 0.016242 mol

So, g per mole

$$\frac{0.7523 \text{ g unknown}}{0.01624 \text{ mol unknown}} = 46.3 \text{ g/mol}$$

Determine the density, in g/cm^3, of CH_4 at 25.0 °C and 991 torr. If we knew the mass and the volume of a sample of gas, it would be simple to calculate the density of the gas.

It is easy to find the volume that 1 mole occupies. Why choose one mole? Because I know the mass of one mole of CH_4!

$$991 \text{ torr} \times \frac{1 \text{ atm}}{760 \text{ torr}} \quad V = (1 \text{ mol})(0.08206 \text{ L} \cdot \text{atm} \cdot \text{K}^{-1} \cdot \text{mol}^{-1})(298.25 \text{ K})$$

18.76 L CH$_4$

$$\frac{16.0426 \text{ g CH}_4}{18.76 \text{ L CH}_4} \times \frac{1 \text{ L CH}_4}{1000 \text{ mL CH}_4} = 8.55 \times 10^{-4} \text{ CH}_4 \text{ g/cm}^3$$

Under the same conditions of temperature and pressure, which gas is more dense, Xe or C$_3$H$_8$?