A balloon is taken from a room where the temperature is 27.0 C to a freezer where the temperature is -5.0 C. If the balloon has a volume of 3.5 L in the 27.0 C room, what is the volume of the balloon in the freezer. Assume pressure is constant.

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

$V_1 = 3.5\,L$ $V_2 = ?$

$T_1 = 27.0°C = 300.2\,K$

$T_2 = -5.0°C = 268.2\,K$

$\frac{3.5\,L}{300.2\,K} = \frac{V_2}{268.2\,K}$

$3.1\,L = V_2$

2.25 L of nitrogen gas is trapped in a piston at 25.0 C and 1.00 atm pressure. If the piston is pushed in so that the gas's volume is 1.00 L while the temperature increases to 31.0 C, what is the pressure of the gas in the piston?

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

$P_1 = 1.00\,atm$; $P_2 = ?$

$V_1 = 2.25\,L$; $V_2 = 1.00\,L$

$T_1 = 25.0°C = 298.2\,K$; $T_2 = 31.0°C = 304.2\,K$

$\frac{(1.00\,atm)(2.25\,L)}{(298.2\,K)} = \frac{P_2(1.00\,L)}{(304.2\,K)}$; $P_2 = 2.30\,atm$

Calculate the mass of 22650* L of oxygen gas at 25.0 C and 1.18 atm pressure.

← O_2

*Volume of a 10'x10'x8' room

O_2: 32.00 g O_2 = mol O_2

1 - Calculate the moles oxygen gas using PV=nRT
2 - Convert moles oxygen gas to mass using FORMULA WEIGHT.

(1) $PV = nRT$ | $P = 1.18$ atm | $R = 0.08206 \frac{L \cdot atm}{mol \cdot K}$

$n = \frac{PV}{RT}$ | $V = 22650$ L | $T = 25.0°C = 298.2$ K

$$n_{O_2} = \frac{(1.18\ atm)(22650\ L)}{\left(0.08206 \frac{L \cdot atm}{mol \cdot K}\right)(298.2\ K)} = 1092.222357\ mol\ O_2$$

(2)

$$1092.222357\ mol\ O_2 \times \frac{32.00\ g\ O_2}{mol\ O_2} = 35\bar{0}00\ g\ O_2$$

35.0 kg

~77 lb

CHEMICAL CALCULATIONS WITH THE GAS LAWS

$FW_{NaHCO_3} = 84.007\ g/mol$

$$H_2SO_4(aq) + 2NaHCO_3(s) \rightarrow 2H_2O(l) + 2CO_2(g) + Na_2SO_4(aq)$$

Given 25.0 g of sodium bicarbonate and sufficient sulfuric acid, what volume of carbon dioxide gas would be produced at 25.0 C and 0.950 atm pressure?

1 - Convert 25.0 g sodium bicarbonate to moles. Use FORMULA WEIGHT.
2 - Convert moles sodium bicarbonate to moles carbon dioxide. Use CHEMICAL EQUATION.
3 - Convert moles carbon dioxide to volume. Use IDEAL GAS EQUATION.

(1) $84.007\ g\ NaHCO_3 = mol\ NaHCO_3$ (2) $2\ mol\ NaHCO_3 = 2\ mol\ CO_2$

$$25.0\ g\ NaHCO_3 \times \frac{mol\ NaHCO_3}{84.007\ g\ NaHCO_3} \times \frac{2\ mol\ CO_2}{2\ mol\ NaHCO_3} = 0.297594248\ mol\ CO_2$$

(1) (2)

(3) $PV = nRT$

$V = \frac{nRT}{P}$

$n = 0.297594248\ mol\ CO_2$ $T = 25.0°C = 298.2\ K$

$R = 0.08206\ \frac{L \cdot atm}{mol \cdot K}$ $P = 0.950\ atm$

$$V = \frac{(0.297594248\ mol\ CO_2)(0.08206\ \frac{L \cdot atm}{mol \cdot K})(298.2\ K)}{(0.950\ atm)} = 7.67\ L\ CO_2$$

at 25.0°C, 0.950 atm

What volume would the gas in the last example problem have at STP?

STP: "Standard Temperature and Pressure" (0 C and 1 atm)

Let's convert to the new set of conditions with the combined gas law!

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

$P_1 = 0.950$ atm $P_2 = 1$ atm

$V_1 = 7.67$ L $V_2 = ?$

$T_1 = 298.2$ K $T_2 = 0°C = 273.2$ K

$$\frac{(0.950 \text{ atm})(7.67 \text{ L})}{(298.2 \text{ K})} = \frac{(1 \text{ atm})V_2}{(273.2 \text{ K})}$$

6.67 L at STP $= V_2$

Alternate solution: Use PV=nRT, with P=1 atm and T = 273.2 K ... since we had already calculated the number of moles 'n' in the previous problem!

$$2\,NH_4NO_3(s) \longrightarrow 2\,N_2(g) + O_2(g) + 4\,H_2O(g)$$

At 300.°C, ammonium nitrate violently decomposes to produce nitrogen gas, oxygen gas, and water vapor. What is the total volume of gas that would be produced at 1.00 atm by the decomposition of 15.0 grams of ammonium nitrate?

To simplify the calculation, we'll calculate the TOTAL MOLES OF GAS instead of the moles of each gas individually

$FW\ NH_4NO_3 = 80.052\ g/mol$

1 - Convert 15.0 g ammonium nitrate to moles. Use FORMULA WEIGHT.
2 - Convert moles ammonium nitrate to TOTAL MOLES GAS. Use CHEMICAL EQUATION
3 - Convert TOTAL MOLES OF GAS to volume using IDEAL GAS EQUATION

(1) $80.052\ g\ NH_4NO_3 = mol\ NH_4NO_3$ (2) $2\ mol\ NH_4NO_3 = 7\ mol\ gas$ $(2+1+4=7)$

$$15.0\ g\ NH_4NO_3 \times \frac{mol\ NH_4NO_3}{80.052\ g\ NH_4NO_3} \times \frac{7\ mol\ gas}{2\ mol\ NH_4NO_3} = 0.6558237146\ mol\ gas$$

(3) $PV = nRT$

$V = \frac{nRT}{P}$

$n = 0.6558237146\ mol\ gas$

$R = 0.08206\ \frac{L \cdot atm}{mol \cdot K}$

$T = 300.°C = 573\ K$

$P = 1.00\ atm$

$$V = \frac{(0.6558237146\ mol\ gas)\left(0.08206\ \frac{L \cdot atm}{mol \cdot K}\right)(573\ K)}{(1.00\ atm)} = 30.8\ L$$

30.8 L at 300°C 1 atm

REAL GASES

- The empirical gas laws (including the ideal gas equation) do not always apply.

  - The gas laws don't apply in situations where the assumptions made by kinetic theory are not valid.

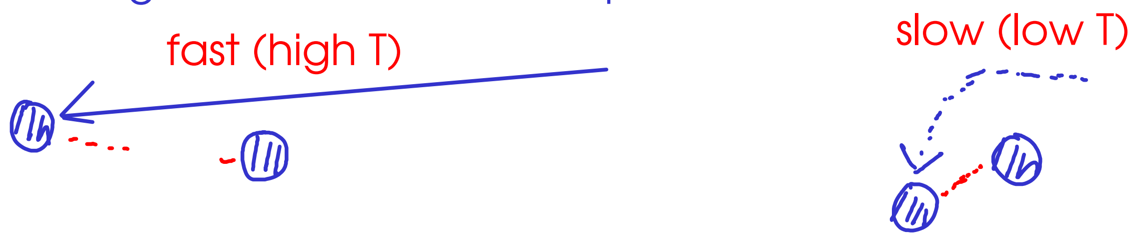
    - When would it be FALSE that the space between gas molecules is much larger than the molecules themselves?

      - at high pressure, molecules would be much closer together!

    - When would it be FALSE that attractive and repulsive forces would be negligible?

      - at high pressure, attractions and repulsions should be stronger!

      - at low temperature, attractions and repulsions have a more significant affect on the paths of molecules



-The gas laws are highly inaccurate near the point where a gas changes to liquid!

- In general, the lower the pressure and the higher the temperature, the more IDEAL a gas behaves.