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.

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2} \xrightarrow{P_1 = P_2} \frac{V_1}{T_1} = \frac{V_2}{T_2}$$

$V_1 = 3.5L$

$T_1 = 27.0°C = 300.2K$

$V_2 = ?$

$T_2 = -5.0°C = 268.2K$

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

$$\boxed{3.1L} = V_2$$

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?

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

$P_1 = 1.00\ atm$ | $P_2 = ?$

$V_1 = 2.25L$ | $V_2 = 1.00L$

$T_1 = 25.0°C = 298.2K$ | $T_2 = 31.0°C = 304.2K$

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

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 - Use the IDEAL GAS EQUATION to find the MOLES of oxygen gas.
2 - Use the FORMULA WEIGHT of oxygen gas to find MASS.

(1) $PV = nRT$

$n = \frac{PV}{RT}$

$P = 1.18\ atm$

$V = 22650\ L$

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

$T = 25.0°C = 298.2\ K$

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

(2) 32.00 g O_2 = mol O_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 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)

$PV = nRT$

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

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

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

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

$= 7.67\ L\ CO_2$ @ $25.0°C$, $0.950\ atm$

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

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

We can change the conditions of the gas using 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

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

$$P_2 = \boxed{6.67\text{ L @ STP}}$$

Alternate solution: We already calculated the moles of gas produced in the experiment. Use the ideal gas equation with that number of moles and the STP conditions to find the volume at STP. (You should get the same answer as above!)

$2\ NH_4NO_3(s) \longrightarrow 2\ N_2(g) + O_2(g) + 4\ H_2O(g)$ $NH_4NO_3 = 80.052\ g/mol$

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 will calculate the TOTAL MOLES OF GAS instead of treating the three gases separately.

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

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

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

(1) (2)

(3) $PV = nRT$

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

$n = 0.6558237146\ mol\ gas$ $P = 1.00\ atm$

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

$T = 300.°C = 573\ K$

$$V = \frac{(0.6558237146\ mol\ gas)\left(0.08206\ \frac{L\cdot atm}{mol\cdot K}\right)(573\ K)}{1.00\ atm} = 30.8\ L\ @\ 300°C,\ 1.00\ atm$$

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.

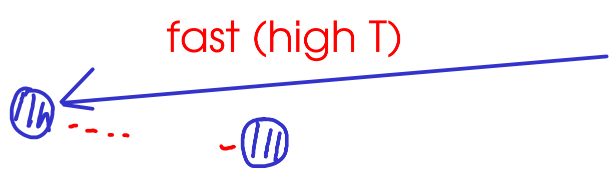
        - When would it be FALSE that the space between gas molecules is much larger than the molecules themselves?

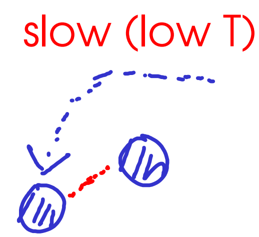
            - at high pressure, molecules would be much closer together!

        - 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.

van der Waals equation

- an attempt to modify PV = nRT to account for several facts.
  - gas molecules actually have SIZE (they take up space)
  - attractive and repulsive forces

$$PV = nRT$$ Ideal gas equation

$$\left(P + \frac{n^2 a}{V^2}\right)(V - nb) = nRT$$ van der Waals equation

nb: attempts to account for molecular size

$\frac{n^2 a}{V^2}$: attempts to account for attractive / repulsive forces

* "a" and "b" are experimentally determined parameters that are different for each gas. p 208

He: a = 0,0346, b = 0,0238 tiny, no special attractive forces

H_2O: a = 5,537, b = 0,03049 small, but strong attractions between moleculres

CH_3CH_2OH: a = 12.56 b = 0,08710 larger, and strong attractions between molecules

2500 L of chlorine gas at 25.0 C and 1.00 atm are used to make hydrochloric acid. How many kilograms of hydrochloric acid could be produced if all the chlorine reacts?

$$H_2 + Cl_2 \rightarrow 2\,HCl$$

1 - Convert 2500 L chlorine gas to moles. Use IDEAL GAS EQUATION.
2 - Convert moles chlorine gas to moles HCl. Use CHEMICAL EQUATION.
3 - Convert moles HCl to mass. Use FORMULA WEIGHT.

(1) $PV = nRT$

$n = \frac{PV}{RT}$

$P = 1.00\text{ atm}$ $\quad R = 0.08206 \frac{L\cdot atm}{mol\cdot K}$

$V = 2500\text{ L}$ $\quad T = 25.0°C = 298.2\text{ K}$

$$n_{Cl_2} = \frac{(1.00\text{ atm})(2500\text{ L})}{(0.08206 \frac{L\cdot atm}{mol\cdot K})(298.2\text{ K})} = 102.164983\text{ mol } Cl_2$$

(2) $mol\ Cl_2 = 2\ mol\ HCl$

(3) $36.458\ g\ HCl = mol\ HCl$

HCl: H: 1 x 1.008
Cl: 1 x 35.45
36.458 g/mol

$$102.164983\text{ mol } Cl_2 \times \frac{2\text{ mol HCl}}{\text{mol } Cl_2} \times \frac{36.458\text{ g HCl}}{\text{mol HCl}} = 7450\text{ g HCl}$$

$kg = 10^3\ g$

$$7450\text{ g HCl} \times \frac{kg}{10^3\ g} = 7.45\text{ kg HCl}$$

$$2\,HCl + Na_2CO_3 \rightarrow CO_2 + H_2O + 2\,NaCl$$

If 48.90 mL of 0.250 M HCl solution reacts with sodium carbonate to produce 50.0 mL of carbon dioxide gas at 290.2 K, what is the pressure of the carbon dioxide gas?

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1 - Convert 48.90 mL of HCl solution to moles. Use MOLARITY.
2 - Convert moles HCl to moles carbon dioxide. Use CHEMICAL EQUATION.
3 - Convert moles carbon dioxide to gas pressure. Use IDEAL GAS EQUATION.

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(1) $0.250 \text{ mol HCl} = L$ $\quad mL = 10^{-3} L$ $\quad$ (2) $2 \text{ mol HCl} = \text{mol } CO_2$

$$48.90 \text{ mL} \times \frac{10^{-3} L}{mL} \times \frac{0.250 \text{ mol HCl}}{L} \times \frac{\text{mol } CO_2}{2 \text{ mol HCl}} = 0.006112S \text{ mol } CO_2$$

(1) $\qquad$ (2)

---

(3) $PV = nRT$

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

$n = 0.006112S \text{ mol } CO_2$ $\quad T = 290.2 K$

$R = 0.08206 \frac{L \cdot atm}{mol\,K}$ $\quad V = 50.0 \text{ mL} = 0.0500 L$

$$P = \frac{(0.006112S \text{ mol } CO_2)\left(0.08206 \frac{L \cdot atm}{mol\,K}\right)(290.2 K)}{(0.0500 L)} = \boxed{2.91 \text{ atm}}$$