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) First, find the MOLES of gas using the ideal gas equation and the information given.
2) Convert moles of gas to mass using formula weight.

$PV = nRT$

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

$P = 1.18\ atm$

$V = 22650\ L$

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

$T = 25.0°C = 298.2\ K$

① $n_{O_2} = \frac{(1.18\ atm)(22650\ L)}{\left(0.08206\ \frac{L \cdot atm}{mol \cdot K}\right)(298.2\ K)} = 1092.222357\ mol\ O_2$

$32.00\,g\ O_2 = mol\ O_2$

② $1092.222357\ mol\ O_2 \times \frac{32.00\,g\ O_2}{mol\ O_2} = \boxed{35\bar{0}00\ g\ O_2}$

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 $2\ mol\ NaHCO_3 = 2\ mol\ CO_2$

$$25.0\ g\ NaHCO_3 \times \underbrace{\frac{mol\ NaHCO_3}{84.007\ g\ NaHCO_3}}_{①} \times \underbrace{\frac{2\ mol\ CO_2}{2\ mol\ NaHCO_3}}_{②} = 0.2975942481\ mol\ CO_2$$

③ $PV = nRT$

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

$n = 0.2975942481\ mol\ CO_2$ $T = 25.0°C = 298.2\ K$

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

$$V = \frac{(0.2975942481\ mol\ CO_2)\left(0.08206\ \frac{L \cdot atm}{mol \cdot K}\right)(298.2\ K)}{(0.950\ atm)} = \boxed{7.67\ L\ CO_2\ \text{at}\ 25.0°C,\ 0.950\ atm}$$

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

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

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

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

$V_1 = 7.67$ L $V_2 =$

$T_1 = 298.2$ K $T_2 = 0°C = 273.2$ K

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

$$\boxed{6.67\ \text{L at STP}} = V_2$$

Alternate solution: Since we already calculated the moles of carbon dioxide, we could plug that and the STP pressure and temperature into the ideal gas equation!

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

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?

Since gas molecules behave the same way under the same conditions, let's simplify the problem by caclulating the TOTAL MOLES GAS instead of treating each gas 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. Use 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 $PV = nRT$

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

$n = 0.6558237146\ mol\ gas$ $T = 300.°C = 573\ K$

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

$$V = \frac{(0.6558237146\ mol\ gas)(0.08206\ \frac{L\cdot atm}{mol\cdot K})(573\ K)}{1.00\ atm} = \boxed{30.8\ L\ gas\ \text{at}\ 300.°C,\ 1.00\ atm}$$

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

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

attempts to account for molecular size

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

$25\bar{0}0$ 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 $PV = nRT$

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

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

$V = 25\bar{0}0\,L$

$T = 25.0°C = 298.2\,K$

$$n_{Cl_2} = \frac{(1.00\text{ atm})(25\bar{0}0\,L)}{\left(0.08206\,\frac{L\cdot atm}{mol\cdot K}\right)(298.2\,K)} = 102.164983\text{ mol } Cl_2$$

2 $\text{mol } Cl_2 = 2\text{ mol HCl}$

3 $HCl - H: 1\times1.008$

$Cl: 1\times35.45$

$36.458\text{ g HCl} = \text{mol HCl}$

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

Report answer in kg!

$kg = 10^3\,g$

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

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

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?

1 - Convert 48.90 mL of 0.250 M HCl to moles. Use MOLARITY.
2 - Convert moles HCl to moles carbon dioxide. Use CHEMICAL EQUATION.
3 - Convert moles carbon dioxide to PRESSURE carbon dioxide. Use IDEAL GAS EQUATION.

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

$$48.90\ mL \times \frac{10^{-3}\,L}{mL} \times \underbrace{\frac{0.250\ mol\ HCl}{L}}_{1} \times \underbrace{\frac{mol\ CO_2}{2\ mol\ HCl}}_{2} = 0.0061125\ mol\ CO_2$$

3 $PV = nRT$

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

$n = 0.0061125\ mol\ CO_2 \quad T = 290.2\ K$

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

$$P = \frac{(0.0061125\ mol\ CO_2)\left(0.08206\ \frac{L \cdot atm}{mol \cdot K}\right)(290.2\ K)}{0.0500\ L} = \boxed{2.91\ atm\ CO_2\ at\ 50.0\ mL,\ 290.2\ K}$$

ENERGY

- thermodynamics: the study of energy transfer

Conservation of energy: Energy may change form, but the overall amount of energy remains constant. "first law of thermodynamics"

- ... but what IS energy?

  - energy is the ability to do "work"

motion of matter

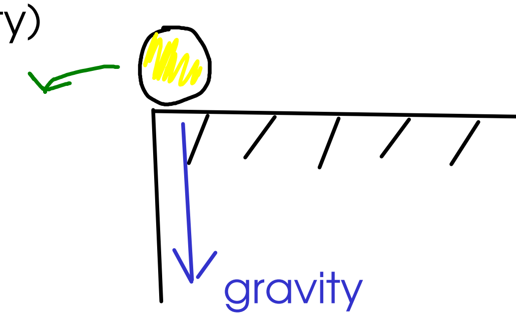
Kinds of energy?

- Kinetic energy: energy of matter in motion $E_K = \frac{1}{2} m v^2$

mass

velocity

- Potential energy: energy of matter that is being acted on by a field of force (like gravity)



When the ball falls, its potential energy is converted to kinetic!