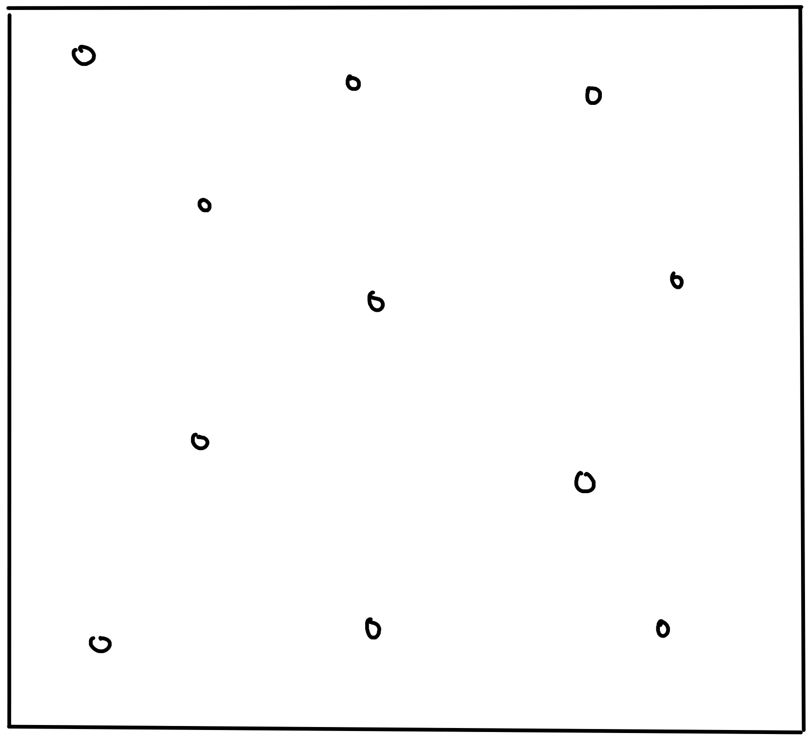
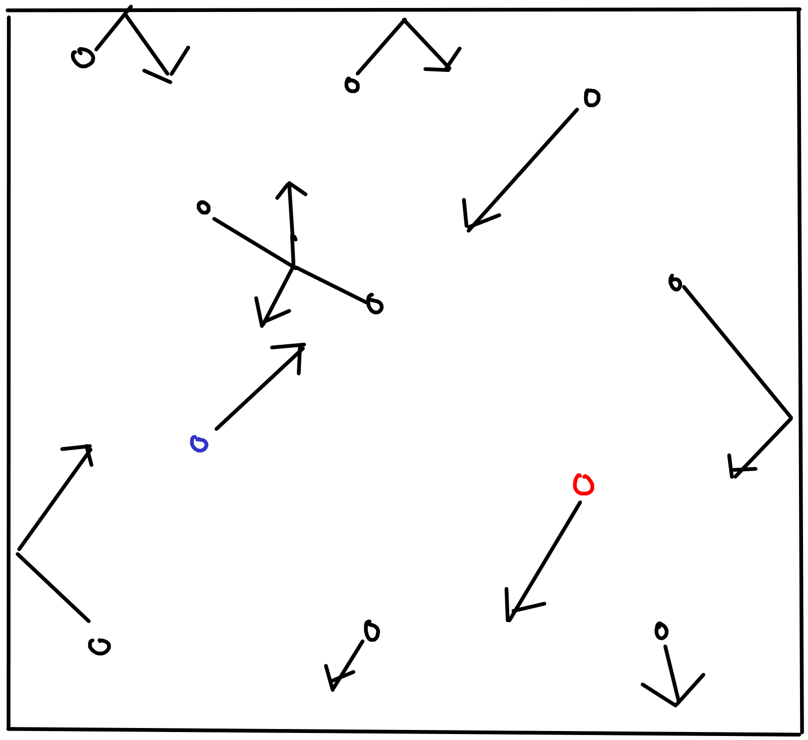
THE KINETIC PICTURE OF GASES



① Gas molecules are small compared to the space between the gas molecules!

LOW DENSITY!



(2) Gas molecules are constantly in motion. They move in straight lines in random directions and with various speeds.

(3) Attractive and repulsive forces between gas molecules are so small that they can be neglected except in a collision.

- Each gas molecule behaves independently of the others.

(4) Collisions between gas molecules and each other or the walls are ELASTIC.

(5) The average kinetic energy of gas molecules is proportional to the absolute temperature.

How does this picture explain the properties of gases?

- Gases expanding to fill their container? Agrees with kinetic picture, since gas molecules are independent
- Thermal expansion of gas at constant pressure? Agrees, because the container has to EXPAND to keep the pressure (from collisions) constant when the gas molecules move faster.
- Pressure increases with temperature at constant volume: Agrees, because the number and force of collisions increases with molecular speed.

GAS LAWS

- were derived by experiment long before kinetic theory, but agree with the kinetic picture!

Boyle's Law:

$$PV = \text{constant}$$

True at constant temperature

$$P_1V_1 = \text{constant} \qquad P_2V_2 = \text{constant}$$

$$\boxed{P_1V_1 = P_2V_2}$$

True at constant temperature

Charles's Law:

$$\frac{V}{T} = \text{constant}$$

True at constant pressure, and using ABSOLUTE temperature

$$\boxed{\frac{V_1}{T_1} = \frac{V_2}{T_2}}$$

True at constant pressure, and using ABSOLUTE temperature

Combined gas law:

$$\frac{PV}{T} = \text{constant}$$

Must use ABSOLUTE temperature units!

↓

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

Must use ABSOLUTE temperature units!

↑ amount (moles) of gas must be constant!

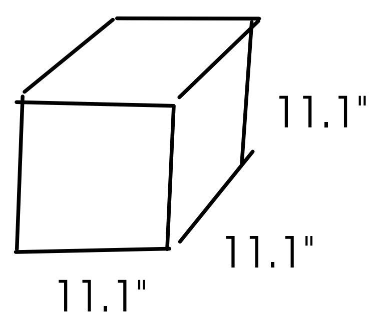
Avogadro's law:

- a mole of any gas at the same conditions has the same volume.

1 mol gas molecules @ 0°C and 1 atm

volume = 22.4 L

"STP"
Standard Temperature and Pressure



= 22.4 L

Ideal gas law:

$$\frac{PV}{T} = \text{constant}$$

... but this constant actually depends on the amount of gas!

$$= n \times \text{"}R\text{"}$$

The ideal gas constant.

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

... combining these together ...

$$\frac{PV}{T} = nR$$

$$PV = nRT$$

P = pressure atm

V = volume L

T = ABSOLUTE temperature K

R = ideal gas constant

n = number of moles of gas molecules

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.

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

$P_1 = \quad P_2 =$ CONSTANT

$V_1 = 3.5 L \quad V_2 = ?$

$T_1 = 27.0°C = 300.2 K \quad T_2 = -5.0°C = 268.2 K$

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

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

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?

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

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

$$\boxed{2.30 atm} = P_2$$

$P_1 = 1.00 atm \quad P_2 = ?$

$V_1 = 2.25 L \quad V_2 = 1.00 L$

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

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 - Convert the oxygen gas's PVT to moles using the ideal gas equation.
2 - Convert moles oxygen gas to mass. Use FORMULA WEIGHT.

$PV = nRT$

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

$P = 1.18$ atm

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

$V = 22650$ L

$T = 25.0°C = 298.2$ K

① $n_{O_2} = \frac{(1.18\ atm)(22650\ L)}{(0.08206 \frac{L \cdot atm}{mol \cdot K})(298.2\ K)} = 1092.22357$ mol O_2

② 1092.22357 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 grams 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 LAW.

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

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

$PV = nRT$

③ $V = \frac{nRT}{P}$

$n = 0.2975942481\ mol\ CO_2$

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

$T = 25.0°C = 298.2\ K$

$P = 0.950\ atm$

$$V = \frac{(0.2975942481\ mol\ CO_2)(0.08206\ \frac{L \cdot atm}{mol \cdot K})(298.2\ K)}{0.950\ atm}$$

$= 7.67\ L$ at $25.0°C$, $0.950\ atm$

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

(0°C, 1 atm)

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

$P_1 = 0.950$ atm

$V_1 = 7.67$ L

$T_1 = 298.2$ K

$P_2 = 1$ atm

$V_2 = ?$

$T_2 = 0°C = 273.2$ 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 $= V_2$

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 molecuIres

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?

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

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

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① $PV = nRT$, $n = \frac{PV}{RT}$

$P = 1.00\text{ atm}$, $V = 2500\text{ L}$, $R = 0.08206\ \frac{\text{L-atm}}{\text{mol}\cdot\text{K}}$, $T = 25.0^\circ C = 298.2\text{ K}$

$$n_{Cl_2} = \frac{(1.00\text{ atm})(2500\text{ L})}{\left(0.08206\ \frac{\text{L-atm}}{\text{mol}\cdot\text{K}}\right)(298.2\text{ K})} = 102.1646983\text{ mol } Cl_2$$

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

③ HCl – H: 1×1.008
Cl: 1×35.45
$36.458\text{ g } HCl = \text{mol } HCl$

$$102.1646983\text{ mol } Cl_2 \times \frac{2\text{ mol } HCl}{\text{mol } Cl_2} \times \frac{36.458\text{ g } HCl}{\text{mol } HCl} = 7450\text{ g } HCl$$

Problem specifies final answer in kg ... so convert.

$\text{Kg} = 10^3\text{ g}$

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

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

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?

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

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

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

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(3) $PV = nRT$

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

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

$R = 0.08206\ \dfrac{\text{L-atm}}{\text{mol}\cdot\text{K}}$ $V = 50.0$ mL $= 0.0500$ L

$$P = \frac{(0.0061125\text{ mol } CO_2)\left(0.08206\ \frac{\text{L-atm}}{\text{mol}\cdot\text{K}}\right)(290.2\text{ K})}{0.0500\text{ L}}$$

$= \boxed{2.91\text{ atm}}$