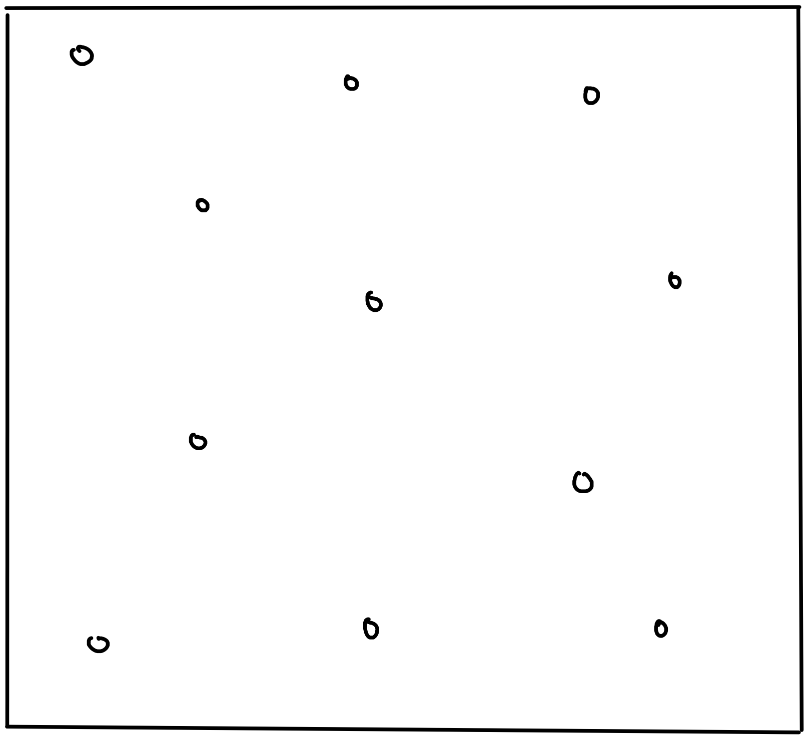
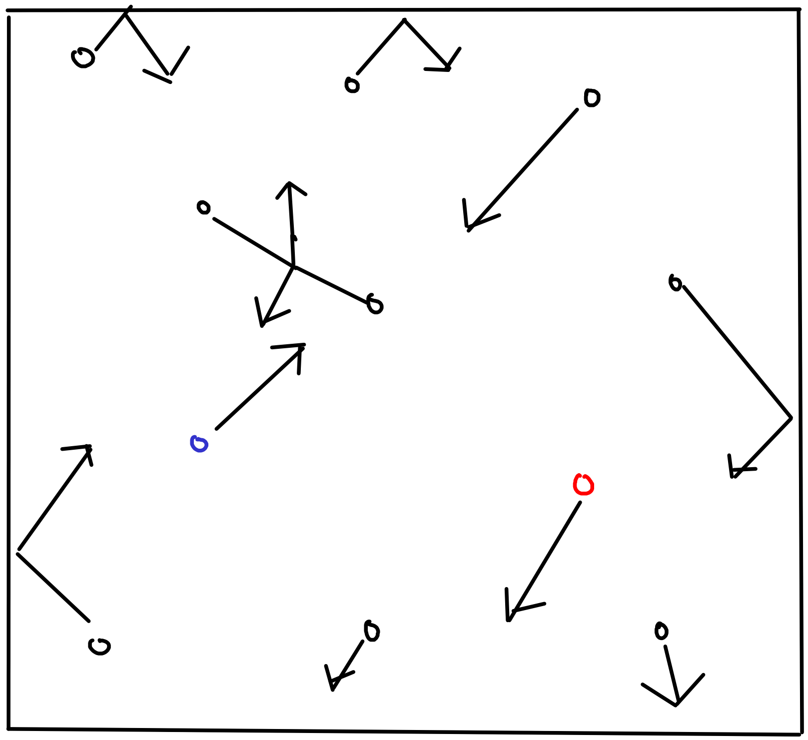
THE KINETIC PICTURE OF GASES



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

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

$\boxed{P_1V_1 = P_2V_2}$ True at constant temperature

Charles's Law: $\dfrac{V}{T} = \text{constant}$ True at constant pressure, and using ABSOLUTE temperature

$\boxed{\dfrac{V_1}{T_1} = \dfrac{V_2}{T_2}}$ True at constant pressure, and using ABSOLUTE temperature

Combined gas law:

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

Must use ABSOLUTE temperature units!

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

Must use ABSOLUTE temperature units!

amount (moles) of gas must be constant!

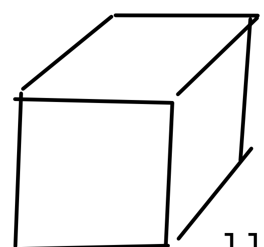
Avogadro's law:

- 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



11.1"

11.1"

11.1"

= 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 ...

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

$$\boxed{PV = nRT}$$

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.

$$\frac{\cancel{P}V_1}{T} = \frac{\cancel{P}V_2}{T_2} \rightarrow \frac{V_1}{T_1} = \frac{V_2}{T_2}$$

$V_1 = 3.5L$

$T_1 = 27.0°C = 300.2K$

$V_2 = ?$

$T_2 = -5.0°C = 268.2K$

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

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

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?

$P_1 = 1.00 \text{ atm}$

$V_1 = 2.25L$

$T_1 = 25.0°C = 298.2K$

$P_2 = ?$

$V_2 = 1.00L$

$T_2 = 31.0°C = 304.2K$

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

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

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

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

O_2

O_2: $\frac{32.00 \text{ g } O_2}{\text{mol } O_2}$

*Volume of a 10'x10'x8' room

1) Use the IDEAL GAS EQUATION, PV=nRT, to find moles of oxygen gas.
2) Convert moles of oxygen gas to mass. Use FORMULA WEIGHT.

① $PV = nRT$

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

$P = 1.18 \text{ atm}$ $V = 22650 \text{ L}$

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

$T = 25.0°C = 298.2 \text{ K}$

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

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

35.0 kg

~77 lb

CHEMICAL CALCULATIONS WITH THE GAS LAWS

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

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

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 doxide. 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 $PV = nRT$

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

$n = 0.2975942481\ mol\ CO_2$ $R = 0.08206\ \frac{L\cdot atm}{mol\cdot K}$

$T = 25.0°C = 298.2\ K$ $P = 0.950\ atm$

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

$\boxed{V = 7.67\ L\ CO_2\ gas}$ at 25.0°C, 0.950 atm

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

(0°C, 1 atm)

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

$P_1 = 0.950$ atm

$V_1 = 7.67$ L

$T_1 = 298.2$ K

$P_2 = 1$ atm

$V_2 = ?$

$T_2 = 0°C = 273.2$ K

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

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

$25\bar{0}0$ L of chlorine gas at 25.0 C and 1.00 atm are used to make hydrochloric acid. How many grams 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 LAW, PV=nRT
2 - Convert moles chlorine gas to moles HCl. Use CHEMICAL EQUATION
3 - Convert moles HCl to mass HCl. Use FORMULA WEIGHT.

---

① $PV = nRT$ | $P = 1.00\text{ atm}$ | $R = 0.08206\,\frac{L\cdot atm}{mol\cdot K}$

$n = \frac{PV}{RT}$ | $V = 25\bar{0}0\text{ L}$ | $T = 25.0°C = 298.2\text{ K}$

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

---

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

③ HCl – H : 1 x 1.008
Cl : 1 x 35.45

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

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

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

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 gas. Use CHEMICAL EQUATION.
3 - Convert moles carbon dioxide gas to pressure. Use IDEAL GAS LAW

① $0.250\text{ mol HCl} = \text{L}$ $\quad \text{mL} = 10^{-3}\text{ L}$ ② $2\text{ mol HCl} = \text{mol CO}_2$

$$48.90\text{ mL} \times \frac{10^{-3}\text{ L}}{\text{mL}} \times \underbrace{\frac{0.250\text{ mol HCl}}{\text{L}}}_{①} \times \underbrace{\frac{\text{mol CO}_2}{2\text{ mol HCl}}}_{②} = 0.006\,1125\text{ mol CO}_2$$

③ $PV = nRT$

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

$n = 0.0061125\text{ mol CO}_2$ $\quad T = 290.2\text{ K}$

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

$$P = \frac{(0.0061125\text{ mol CO}_2)\left(0.08206\ \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(290.2\text{ K})}{0.0500\text{ L}}$$

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