Gas Laws

- Gases change volume with changes in temperature and pressure.
- Gas laws are math equations that allow us to predict these changes.
- Gas laws work well with gases but **DO NOT** work with liquids or solids.
- Gases near their boiling points exhibit deviations from the gas laws because of the vanderWaals forces.

Ideal Gas

- An ideal gas is one that obeys all the gas laws at all temperatures and pressures.
- This means that an ideal gas never becomes a liquid (or solid) and has no vanderWaals forces.
- Of course ideal gases do not actually exist but real gases follow the gas laws very well if they are not near the boiling (or condensation) point.

Boyle's Law

- The pressure of a gas is inversely proportional to the volume if the amount of gas and temperature are held constant.
- In math form $P \propto 1/V$ or
- $PV = \text{constant}$ or
- $P_1V_1 = P_2V_2$
- P is pressure (atm, mm Hg, kPa, torr)
- V is volume (L, mL, cc, cm^3, dm^3)

Charles' Law

- The volume of a gas is directly proportional to the temperature if the pressure and amount of the gas remain constant.
- In math form $V \propto T$ or

$$\frac{V}{T} = \text{constant or} \qquad \frac{V_1}{T_1} = \frac{V_2}{T_2}$$

V is volume and

T is temperature **in Kelvin**

$K = °C + 273.15$

Gay-Lussac's Law

- The pressure of a gas is directly proportional to the temperature if the amount of gas and volume remain constant.
- In math form $P \propto T$ or

$$\frac{P}{T} = \text{constant or} \qquad \frac{P_1}{T_1} = \frac{P_2}{T_2}$$

Combined Gas Law

Combining all three into one equation (so we only have to remember one equation).

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

Pressure and volume can be in any units as long as they match. Temperature must be in Kelvin so there is no negative numbers.

Gas Law Standards

Standard Temperature and Pressure - STP

Standard temperature is 0 °C (273.15 K)

Standard pressure is 1.000 atmosphere

1 atm = 101.3 kPa = 760 mm Hg = 760 torr

This is because 0 °C is easy to get (ice water) and 1.000 atmosphere is "normal" at sea level.

At STP, one mole of any gas =22.4L

Gas Law Problems

35.6 mL of a gas is collected at 83.42 kPa and 23.0 °C. What will be the volume at 34.37 kPa and 85.3 °C?

Step 1, make a variable list and fill it in.

$V_1 = 35.6$ mL

$P_1 = 83.42$ kPa

$T_1 = 23.0$ °C $= 296.15$ K

$V_2 = ?$

$P_2 = 34.37$ kPa

$T_2 = 85.3$ °C $= 358.45$ K

Problem Solution - solve for the unknown quantity

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

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

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

Substitute the values from the variable list into the equation.

$$\frac{(83.42\ \cancel{kPa})(35.6\ mL)(358.45\ \cancel{K})}{(296.15\ \cancel{K})(34.37\ \cancel{kPa})} = V_2 = 104.6\ mL$$

Ideal Gas Law

$PV = nRT$

P is pressure (kPa)

V is volume (liters)

n is number of moles

R is the universal gas constant

$$R = \frac{8.314\ \text{L kPa}}{\text{mole K}}$$

T is temperature in Kelvin

Ideal Gas Law

The Ideal Gas Law allows us to do stoichiometry because using the Ideal Gas Law it is possible to calculate moles of gases.

Moles of gases only, not solids, not liquids, not solutions (like when we have molarity and mL).

Ideal Gas Law

What volume of oxygen is required to burn 45.23 grams of propane (C_3H_8) at 24.6°C and 97.2 kPa?

$$C_3H_{8\,(g)} + 5O_{2\,(g)} \rightarrow 3CO_{2\,(g)} + 4H_2O_{(g)}$$

$$45.23\text{ g }C_3H_8 \times \frac{1\text{ mole }C_3H_8}{44.0932\text{ g }C_3H_8} \times \frac{5\text{ mole }O_2}{1\text{ mole }C_3H_8}$$

$$= 5.1289\text{ moles }O_2$$

Ideal Gas Law

$n = 5.1289$ moles O_2

$P = 97.2$ kPa

$V = ?$

$T = 24.6°C = 297.75$ K

$R = 8.314$ L kPa/ mole K

$PV = nRT$ $\qquad$ $V = nRT/P$

$$V = \frac{(5.1289\ \cancel{\text{moles}})\left(\dfrac{8.314\ \text{L}\ \cancel{\text{kPa}}}{\cancel{\text{mole}}\ \cancel{\text{K}}}\right)(297.75\ \cancel{\text{K}})}{97.2\ \cancel{\text{kPa}}} = 131\ \text{L}$$

Dalton's Law of Partial Pressure

The total pressure of a gas is equal to the sum of the partial pressures.

$P_T = P_1 + P_2 + P_3 + \ldots$

This law is most commonly used in high school chemistry when a gas is collected over water.

To correctly calculate the amount of gas we must mathematically remove the water.

Dalton's Law of Partial Pressure

35.26 mL of a gas is **collected over water** at 94.26 kPa and 54.0 °C. What will be the volume of the dry gas at STP?

$P_T = P_{gas} + P_{water} = 94.26$ kPa

$P_{water} = 15.012$ kPa

$P_{gas} = P_T - P_{water} = 94.26$ kPa $- 15.012$ kPa

$P_{gas} = 79.248$ kPa

$P_1 = 79.248$ kPa

$V_1 = 35.26$ mL

$T_1 = 54.0$ °C $= 327.15$ K

$P_2 = 101.3$ kPa

$T_2 = 0$ °C $= 273.15$ K

$V_2 = ?$

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

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

$$\frac{P_1V_1T_2}{T_1P_2} = \frac{V_2}{}$$

$$\frac{(79.248\ \cancel{\text{kPa}})(35.26\ \text{mL})(273.15\ \cancel{\text{K}})}{(327.15\ \cancel{\text{K}})(101.3\ \cancel{\text{kPa}})} = V_2 = 23.04\ \text{mL}$$

Ideal Gas Law

Special Situation - at STP - what is the volume of one mole of a gas at STP?

n = 1 mole

T = 273.15 K

P = 101.3 kPa

V = ?

$$V = \frac{(1.000\ \cancel{\text{moles}})\left(\dfrac{8.314\ \text{L}\ \cancel{\text{kPa}}}{\cancel{\text{mole}}\ \cancel{\text{K}}}\right)(273.15\ \cancel{\text{K}})}{101.3\ \cancel{\text{kPa}}} = 22.41\ \text{L}$$

Ideal Gas Law

We can use the shortcut that one mole of any gas at STP is equal to 22.41 L.

We could do this for ANY conditions.

This is an effect of the Ideal Gas Law from Avogradro → Equal volumes of gas at equal temperature and pressure contain equal numbers of molecules.

This relationship can be used to simplify some calculations. It is NOT required.

Avogradro's Hypothesis

What volume of oxygen is required to burn 43.2 L of methane at STP?

$$CH_{4\,(g)} + 2O_{2\,(g)} \rightarrow CO_{2\,(g)} + 2H_2O_{\,(g)}$$

$$43.2\text{ L }CH_{4\,(g)} \times \frac{1\text{ mole }CH_4}{22.41\text{ L }CH_4} \times \frac{2\text{ mole }O_2}{1\text{ mole }CH_4} \times \frac{22.41\text{ L }O_2}{1\text{ mole }O_2}$$

86.4 L O_2

Grahams Law of Diffusion

- Gasses have high rates of diffusion.
- The rate of diffusion depends on how fast they are moving.
- The velocity of the molecules is related to the kinetic energy.
- We measure the kinetic energy of molecules by measuring temperature.
- $K = 1/2\ mv^2$

Grahams Law of Diffusion

For two gases in the same container the temperature must be the same → $K_1 = K_2$

$1/2\ m_1 v_1^2 = 1/2\ m_2\ v_2^2$

$m_1\ v_1^2 = m_2\ v_2^2$

$$\frac{m_1}{m_2} = \frac{v_2^2}{v_1^2} \qquad\qquad \frac{\sqrt{m_1}}{\sqrt{m_2}} = \frac{v_2}{v_1}$$

Grahams Law of Diffusion

How fast does helium diffuse compared to nitrogen?

(find the ratio of v_{He} to $v_{nitrogen}$)

m_{helium} = 4.0026 g/mole

$m_{nitrogen}$ = 28.01348 g/mole

$$\frac{\sqrt{m_1}}{\sqrt{m_2}} = \frac{v_2}{v_1} \qquad \frac{\sqrt{28.01348\ \text{g/mole}}}{\sqrt{4.0026\ \text{g/mole}}} = \frac{v_2}{v_1}$$

2.645 helium : 1.000 nitrogen