

Combined Gas Law

The Combined Gas Law

- Combining Boyle's Law and Charles' Law results in a new gas law, the **Combined Gas Law**, which is used when pressure, volume and temperature all change while the number of gas molecules remain constant.

$$\frac{PV}{T} = k \quad \text{therefore} \quad \frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

The Combined Gas Law

- Example:**
Find the volume of a gas (V_2) at 110.0 kPa (P_2) and 35.0°C (T_2) if its volume at 98.0 kPa (P_1) and 15.0 0°C (T_1) is 7.50 L (V_1)?

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2} \quad \text{rearranges to} \quad V_2 = \frac{P_1V_1T_2}{T_1P_2}$$

$$V_2 = \frac{(98.0 \text{ kPa})(7.50 \text{ L})(308.2 \text{ K})}{(110.0 \text{ kPa})(288.2 \text{ K})} = 7.1 \text{ L}$$

Avogadro's Hypothesis

- Equal volumes of gases at the same temperature and pressure contain the same number of molecules.



1.0 L H₂



1.0 L He

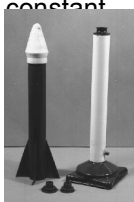


1.0 L Cl₂

At 0°C and 101.325 kPa, 1 L of each of these gases contains 2.7×10^{22} molecules

Avogadro's Hypothesis

- The volume of a gas depends on the number of molecules in the container as long as the temperature and pressure are constant.



Avogadro's Hypothesis

- Avogadro observed the following:

1 volume of hydrogen + 1 volume of chlorine \longrightarrow 2 volumes of hydrogen chloride

- For this to be true, hydrogen, chlorine and hydrogen chloride must be **diatomic molecules**. (Why?)

No.

No.

Yes!

Calculations Involving Molar Volume at STP

- If Avogadro's Hypothesis is correct, it should be possible to obtain a definite number of gaseous molecules by measuring a definite volume of gas.
- One mole of almost any common gas at 0°C and 101.325 kPa (STP – *standard temperature and pressure*) is approximately **22.4 L/mol**.

Calculations Involving Molar Volume at STP

Equation Method

$$n_{\text{any gas}} = \frac{V}{V} \quad \text{or} \quad n_{\text{any gas}} = \frac{V}{22.4 \text{ L/mol}}$$

$$V = n_{\text{any gas}} \times V \quad \text{or} \quad V = n_{\text{any gas}} \times 22.4 \text{ L/mol}$$

Conversion Factor Method

$$n_{\text{any gas}} = V \times \frac{1 \text{ mol}}{22.4 \text{ L}}$$

$$V = n_{\text{any gas}} \times \frac{22.4 \text{ L}}{1 \text{ mol}}$$

Calculations Involving Molar Volume at STP

Example #1	Solving Process	Equation Method
Calculate the number of moles of oxygen gas contained in a volume of 112 L at STP.	$n_{\text{O}_2} = \frac{V}{V} = \frac{112 \text{ L}}{22.4 \text{ L/mol}} = 5.00 \text{ mol O}_2$	$V = \frac{m}{d}$ $d = \frac{m}{V}$ $V = m \times \frac{1}{d}$
	$112 \text{ L O}_2 \times \frac{1 \text{ mol O}_2}{22.4 \text{ L O}_2} = 5.00 \text{ mol O}_2$	
Example #2	Solving Process	
Calculate the volume (at STP) of hydrogen gas produced at STP.	$n_{\text{H}_2} = 0.200 \text{ mol} \times 22.4 \text{ L/mol} = 4.48 \text{ L}$	
	$0.200 \text{ mol H}_2 \times \frac{22.4 \text{ L}}{1 \text{ mol H}_2} = 4.48 \text{ L H}_2$	
Example #3	Solving Process	Conversion Factor Method
Calculate the mass of oxygen gas present in a volume of 56.0 L at STP.	$n_{\text{O}_2} = \frac{V}{V} = \frac{56.0 \text{ L}}{22.4 \text{ L/mol}} = 2.50 \text{ mol}$ $m_{\text{O}_2} = n \times M = 2.50 \text{ mol} \times 32.00 \text{ g/mol} = 80.0 \text{ g}$	$\text{volume of any substance} \xrightarrow{\text{density}} \text{mass}$ $\text{mass} \xrightarrow{\text{molar mass}} \text{moles}$ $\text{moles} \xrightarrow{\text{molar volume}} \text{Volume of Gas}$
	$56.0 \text{ L O}_2 \times \frac{1 \text{ mol O}_2}{22.4 \text{ L O}_2} = 2.50 \text{ mol O}_2$ $2.50 \text{ mol O}_2 \times \frac{32.00 \text{ g O}_2}{1 \text{ mol O}_2} = 80.0 \text{ g}$	
Example #4	Solving Process	
Calculate the volume occupied at STP by 45.0 g of methane.	$n_{\text{CH}_4} = \frac{m}{M} = \frac{45.0 \text{ g}}{16.04 \text{ g/mol}} = 2.81 \text{ mol}$ $V_{\text{CH}_4} = n \times V = 2.81 \text{ mol} \times 22.4 \text{ L/mol} = 62.9 \text{ L}$	
	$45.0 \text{ g CH}_4 \times \frac{1 \text{ mol CH}_4}{16.04 \text{ g CH}_4} = 2.81 \text{ mol CH}_4$ $2.81 \text{ mol CH}_4 \times \frac{22.4 \text{ L}}{1 \text{ mol CH}_4} = 62.9 \text{ L CH}_4$	

The Ideal Gas Law

- **Boyle's Law:** $V = k_1 \left(\frac{1}{P} \right)$
- **Charles' Law:** $V = k_2 T$
- **Avogadro's Law:** $V = k_3 n$
- **Ideal Gas Law:** $V = \frac{k_1 k_2 k_3 n T}{P}$

OR $V = \frac{R n T}{P}$ where R replaces $k_1 k_2 k_3$

OR $PV = nRT$

The Ideal Gas Law

- The greatest advantage of the **Ideal Gas Law** is that it can be used to calculate the volume, pressure, temperature, or number of moles of gas present at conditions other than 0°C and 101.325 kPa.

where: P = pressure (kPa)
V = volume in liters (L)
n = number of moles of gas (mol)
R = universal gas constant (mol·K)
T = absolute temperature (K)

Last slide for today The Ideal Gas Law

- **Example:** What mass of hydrogen is produced if the gas occupies a volume of 50.0 mL at 20.00°C and 94.01 kPa?

The Ideal Gas Law

Step 1: Calculate the number of moles of H₂.

$$n_{\text{H}_2} = \frac{PV}{RT} = \frac{(94.01 \text{ kPa})(0.0500 \text{ L})}{(8.314 \text{ kPa} \cdot \text{L} \cdot \text{mol}^{-1} \cdot \text{K}^{-1})} = 0.00193 \text{ mol}$$

Step 2: Convert number of moles to mass.

$$m_{\text{H}_2} = nM = 0.00193 \text{ mol H}_2 \times \frac{2.02 \text{ g H}_2}{1 \text{ mol H}_2} = 0.00390 \text{ g H}_2$$
