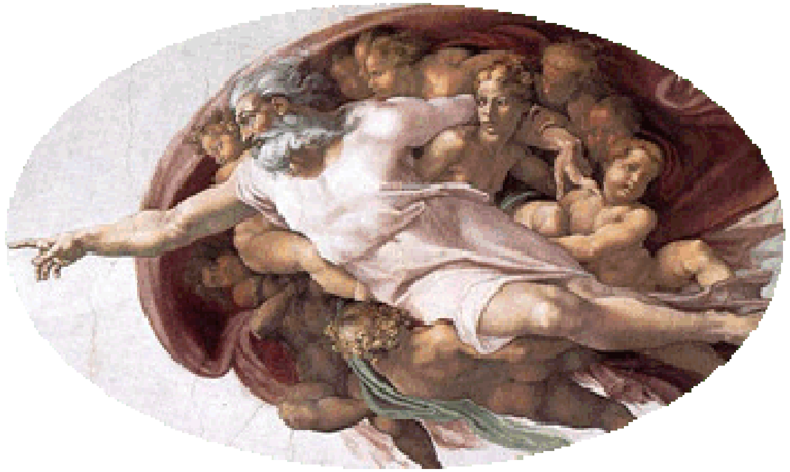
The Ideal Gas Law

$$PV = nRT$$



Ideal Gases

An "ideal" gas exhibits certain theoretical properties. Specifically, an ideal gas …

- Obeys all of the gas laws under all conditions.
- Does not condense into a liquid when cooled.
- Shows perfectly straight lines when its V and T & P and T relationships are plotted on a graph.

In reality, there are no gases that fit this definition perfectly. We assume that gases are ideal to simplify our calculations.

We have done calculations using several gas laws (Boyle's Law, Charles's Law, Combined Gas Law). There is one more to know…

The Ideal Gas Law

$$PV = nRT$$

P = Pressure (in kPa)   V = Volume (in L)

T = Temperature (in K)   n = moles

$$R = 8.31 \frac{kPa \bullet L}{K \bullet mol}$$

R is constant. If we are given three of P, V, n, or T, we can solve for the unknown value.

Recall, From Boyle's Law:

$P_1V_1 = P_2V_2$ or PV = constant

From combined gas law:

$P_1V_1/T_1 = P_2V_2/T_2$ or PV/T = constant

Developing the ideal gas law equation

PV/T = constant. What is the constant?

At STP: T= 273K, P= 101.3 kPa, V= 22.4 L**/mol**

Because V depends on mol, we can change equation to: $\frac{PV}{T \bullet mol}$ = constant

Mol is represented by n, constant by R: $\frac{PV}{Tn} = R$

Rearranging, we get: $PV = nRT$

At STP: (101.3 kPa)(22.4 L) = (1 mol)(R)(273K)

$R = 8.31 \frac{kPa \bullet L}{K \bullet mol}$

Note: always use kPa, L, K, and mol in ideal gas law questions (so units cancel)

Sample problems

How many moles of H_2 is in a 3.1 L sample of H_2 measured at 300 kPa and 20°C?

PV = nRT   P = 300 kPa, V = 3.1 L, T = 293 K

(300 kPa)(3.1 L) = n (8.31 kPa•L/K•mol)(293 K)

$$\frac{(300\ \text{kPa})(3.1\ \text{L})}{(8.31\ \text{kPa}\bullet\text{L/K}\bullet\text{mol})(293\ \text{K})} = n = 0.38\ \text{mol}$$

How many grams of O_2 are in a 315 mL container that has a pressure of 12 atm at 25°C?

PV = nRT P= 1215.9 kPa, V= 0.315 L, T= 298 K

$$\frac{(1215.9\ \text{kPa})(0.315\ \text{L})}{(8.31\ \text{kPa}\bullet\text{L/K}\bullet\text{mol})(298\ \text{K})} = n = 0.15\underline{4}7\ \text{mol}$$

0.15$\underline{4}$7 mol x 32 g/mol = 4.95 g

Ideal Gas Law Questions

1. How many moles of CO_2(g) is in a 5.6 L sample of CO_2 measured at STP?
2. a) Calculate the volume of 4.50 mol of SO_2(g) measured at STP. b) What volume would this occupy at 25℃ and 150 kPa? (solve this 2 ways)
3. How many grams of Cl_2(g) can be stored in a 10.0 L container at 1000 kPa and 30℃?
4. At 150℃ and 100 kPa, 1.00 L of a compound has a mass of 2.506 g. Calculate its molar mass.
5. 98 mL of an unknown gas weighs 0.087 g at SATP. Calculate the molar mass of the gas. Can you determine the identity of this unknown gas?

1. Moles of CO_2 is in a 5.6 L at STP?

P=101.325 kPa, V=5.6 L, T=273 K PV = nRT

(101.3 kPa)(5.6 L) = n (8.31 kPa•L/K•mol)(273 K)

$$\frac{(101.325\text{ kPa})(5.6\text{ L})}{(8.31\text{ kPa}\bullet\text{L/K}\bullet\text{mol})(273\text{ K})} = n = 0.25\text{ mol}$$

2. a) Volume of 4.50 mol of SO_2 at STP.

P= 101.3 kPa, n= 4.50 mol, T= 273 K   PV=nRT

(101.3 kPa)(V)=(4.5 mol)(8.31 kPa•L/K•mol)(273 K)

$$V = \frac{(4.50\text{ mol})(8.31\text{ kPa}\bullet\text{L/K}\bullet\text{mol})(273\text{ K})}{(101.3\text{ kPa})} = 10\underline{0}.8\text{ L}$$

2. b) Volume at 25℃ and 150 kPa (two ways)?

Given: P = 150 kPa, n = 4.50 mol, T = 298 K

$$V = \frac{(4.50\text{ mol})(8.31\text{ kPa}\bullet\text{L/K}\bullet\text{mol})(298\text{ K})}{(150\text{ kPa})} = 74.3\text{ L}$$

From a): P = 101.3 kPa, V = 10<u>0</u>.8 L, T = 273 K

Now P = 150 kPa, V = ?, T = 298 K

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

$$\frac{(101.3\text{ kPa})(100\text{ L})}{(273\text{ K})} = \frac{(150\text{ kPa})(V_2)}{(298\text{ K})}$$

$$(V_2) = \frac{(101.3\text{ kPa})(10\underline{0}.8\text{ L})(298\text{ K})}{(273\text{ K})(150\text{ kPa})} = 74.3\text{ L}$$

3. How many grams of Cl_2(g) can be stored in a 10.0 L container at 1000 kPa and 30℃?

PV = nRT   P= 1000 kPa, V= 10.0 L, T= 303 K

$$\frac{(1000\ \text{kPa})(10.0\ \text{L})}{(8.31\ \text{kPa}\bullet\text{L/K}\bullet\text{mol})(303\ \text{K})} = n = 3.97\ \text{mol}$$

3.97 mol x 70.9 g/mol = 282 g

4. At 150℃ and 100 kPa, 1.00 L of a compound has a mass of 2.506 g. Calculate molar mass.

PV = nRT   P= 100 kPa, V= 1.00 L, T= 423 K

$$\frac{(100\ \text{kPa})(1.00\ \text{L})}{(8.31\ \text{kPa}\bullet\text{L/K}\bullet\text{mol})(423\ \text{K})} = n = 0.028\underline{4}5\ \text{mol}$$

g/mol = 2.506 g / 0.028<u>4</u>5 mol = 88.1 g/mol

5. 98 mL of an unknown gas weighs 0.081 g at SATP. Calculate the molar mass.

PV = nRT P= 100 kPa, V= 0.098 L, T= 298 K

$$\frac{(100\ \text{kPa})(0.098\ \text{L})}{(8.31\ \text{kPa}\bullet\text{L/K}\bullet\text{mol})(298\ \text{K})} = n = 0.003\underline{9}6\ \text{mol}$$

$$\text{g/mol} = 0.081\ \text{g} / 0.003\underline{9}6\ \text{mol} = 2\underline{0}.47\ \text{g/mol}$$

It’s probably neon

(neon has a molar mass of 20.18 g/mol)

Determining the molar mass of butane

Using a butane lighter, balance, and graduated cylinder determine the molar mass of butane.

- Determine the mass of butane used by weighing the lighter before and after use.
- The biggest source of error is the mass of H_2O remaining on the lighter. As a precaution, dunk the lighter & dry well before measuring initial mass. After use, dry well before taking final mass. (Be careful not to lose mass when drying).
- When you collect the gas, ensure no gas escapes & that the volume is 90 – 100 mL.
- Place used butane directly into fume hood.
- Submit values for mass, volume, & g/mol.

Molar Mass of Butane: Data & Calculations

Atmospheric pressure: Temperature: