Kinetic Molecular Theory

Through KMT, several Laws were developed to help calculate the changes in **pressure**, **temperature**, and **volume** of gases.

There are 6 Basic gas Laws:

1. Boyle's Law
2. Charles' Law
3. Gay-Lussac's Law

} Combined Gas Law

4. Avogadro's Law
5. Ideal Gas Law – volume liters only
6. Dalton's Law

v. Avogadro's law

- In 1811, Amedeo Avogadro (an Italian scientist) suggested that at the same temperature and pressure, equal volumes of gases contain equal numbers of moles.
- This observation is called Avogadro's law, shown as:

$$V \propto n$$

$$\frac{V}{n} = k$$

- This equation states that for a gas at constant temperature and pressure, the volume is directly proportional to the number of moles (molecules) of gas.
- In other words, equal numbers of molecules of different gases compared at the same temperature and pressure occupy equal volumes.

- For two samples of gas at the same temperature and pressure, the relation between volumes and numbers of moles can be represented as:

$$\frac{V_1}{n_1} = \frac{V_2}{n_2}$$

- The volume occupied by a mole of gas at STP, is called the standard molar volume.
- It is nearly constant for all gases.
- The standard molar volume of an ideal gas is equal to 22.4 L per mole at STP.
- For an ideal gas: 1mol gas = 22.4 L gas

Example 3.5

Suppose we have a 12.2 L sample containing 0.50 moles oxygen gas (O_2) at a pressure of 1 atm and a temperature of 25 °C. If all this O_2 were converted to ozone (O_3) at the same temperature and pressure, what would be the volume of the ozone?

Solution: The balanced equation for the reaction is:

$$3O_2(g) \rightarrow 2O_3(g)$$

To calculate the moles of O_3 produced, we must use the appropriate mole ratio:

$$\text{Mol of } O_3(g) = 0.5 \cancel{\text{ mol } O_2} \times \frac{2 \text{ mol } O_3}{3 \cancel{\text{mol} O_2}} = 0.33 \text{ mol}$$

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Then, to determine the olume of 0.33 mol O_3, use Equation 3.15.

$$\frac{V_1}{n_1} = \frac{V_2}{n_2}$$

Where V_1 is the volume of n_1 moles of O_2 gas and V_2 is the volume of n_2 moles of O_3 gas. In this case we have

$n_1 = 0.5$ mol $\qquad n_2 = 0.33$ mol

$V_1 = 12.2$ L $\qquad V_2 = ?$

$$\text{Thus, } V_2 = \frac{V_1 n_2}{n_1} = \frac{12.2 \text{ L} \times 0.33 \cancel{\text{ mol}}}{05 \cancel{\text{ mol}}} = 8.052 L = 8.0 L$$



Exercise 3.8

Avogadro's Law

Avogadro's Law – equal volumes of gases at the same temperature and pressure contain equal numbers of molecules.



1 mole of ANY gas takes up a volume of 22.4 L at **STP**. This is called **Molar Volume**

22.4L = 1 mole of gas at STP

vi. The Ideal Gas Equation

- An ideal gas is a hypothetical gas that obeys the gas laws.
- Real gases only obey the ideal gas laws closely at high temperature and low pressure.
- The ideal gas law is a combination of Boyle's law, Charles' law and Avogadro's law.

Boyle's law: $V \alpha P$(at constant T and n)

Charles' law: $V\alpha T$(at constant P and n)

Avogadro's law: $V \alpha n$(at constant P and T)

Example 3.6

What pressure, in atm, is exerted by 54.0 g of Xe in a 1.00 L flask at 20.0 °C? (Molar mass of Xe = 131.3 g/mol)

Solution: By rearranging Equation 3.16

$$P = Rn\frac{T}{V}$$

But n = m/M, Where m is given mass and M is molar mass of a substance Thus,

$$P = R\frac{mT}{MV}$$

$$V = \frac{0.082\ \text{atm}\,\cancel{\text{L}}/\cancel{\text{molK}} \times 54\,\cancel{\text{g}} \times 193\,\cancel{\text{K}}}{131.3\frac{\cancel{\text{g}}}{\cancel{\text{mol}}} \times 1\,\cancel{\text{L}}} = 9.88\ \text{atm}$$

At STP, 0.280 L of a gas weighs 0.400 g. Calculate the molar mass of the gas.

Solution:

Given: V = 0.280 L, m = 0.400 g, T = 273 K and P = 1 atm

So to solve the molar mass of the gas, we use the following formula:

$$M = R\frac{mT}{VP}$$

$$0.082\frac{\cancel{\text{atm L}}}{\text{mol}\,\cancel{\text{K}}} \times \frac{0.400\ \text{g} \times 273\,\cancel{\text{K}}}{0.280\,\cancel{\text{L}} \times 1\,\cancel{\text{atm}}} = 31.98\frac{g}{mol}$$

Exercise 3.9

vii. Graham's Law of Diffusion

- We have discussed that the gas molecules are in constant, rapid, random motion and occupy quickly throughout any container. This spreading of gas molecules throughout the container is called **diffusion.**

- Thomas Graham (1805-869), a Scottish chemist, studied the rate of diffusion of different gases.
- He found that gases having low densities diffuse faster than gases that have higher densities.
- Based on his observations, in 1829, he postulated the relationship between the density of a gas and its rate of diffusion, which is known as **Graham's law of diffusion.**

For two gases (Gas 1 and Gas 2), their rates of diffusion can be given as:

$$r \ \alpha \ \sqrt{\frac{1}{d_1}} \quad Or \quad r \ \alpha \ \sqrt{\frac{1}{M_1}} \tag{3.19}$$

$$r \ \alpha \ \sqrt{\frac{1}{d_2}} \quad \text{Or} \quad r \ \alpha \ \sqrt{\frac{1}{M_2}} \tag{3.20}$$

By rearranging Equations 3.19 and 3.20 we obtainequation 3.21:

$$\frac{r_1}{r_2} = \sqrt{\frac{d_2}{d_1}} \quad Or \quad \frac{r_1}{r_2} = \sqrt{\frac{M_2}{M_1}} \tag{3.21}$$

- **Graham's law of diffusion states that, at constant temperature and pressure, the rate of diffusion of a gas, r, is inversely proportional to the square root of its density, d, or molar mass, M.**

Mathematically it is expressed as:

$$r \propto \sqrt{\frac{1}{d}} \qquad Or \qquad r \propto \sqrt{\frac{1}{M}}$$

- **This relationship indicates how the volume of gas depends on pressure, temperature and number of moles.**

$$V \propto n\frac{T}{P}$$

$$V = Rn\frac{T}{P}$$

At STP ($n = 1$ mol, $V = 22.4$ L, $T = 273$ K (0 °C) and $P = 1$ atm), the values of R can be calculated from the ideal gas equation:

$$R = \frac{PV}{nT} = \frac{1\text{ atm} \times 22.4\text{ L}}{1\text{ mol} \times 273\text{ K}} = 0.082\text{ atm.L/mol.K}$$

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Graham's Law of diffusion

$$\frac{r_1}{r_2} = \sqrt{\frac{u_1^2}{u_2^2}} \quad \text{——} \quad 1$$

From the Kinetic gas equation, $u_1^2 = \frac{3RT}{M_1}$ and $u_2^2 = \frac{3RT}{M_2}$

where M_1 and M_2 = molar mass of two gases

Putting the value u_1^2 and u_2^2 in equation 1

$$\frac{r_1}{r_2} = \sqrt{\frac{M_2}{M_1}}$$

It is the Graham's Law of diffusion

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Laws	Boyle's Law	Charles law	Gay-Lussac Law
Statement	The absolute pressure of a given mass of a perfect gas varies inversely as its volume, when the temperature remains constant	1. The volume of a given mass of a perfect gas varies directly as its absolute temperature, when the absolute pressure remains constant 2. All perfect gases change in volume by 1/273 th of its original volume at 0 °C for every 1 ^{0}C change in temperature, when the pressure remains constant.	The absolute pressure of a given mass of a perfect gas varies directly as its absolute temperature, when the volume remains constant.
Mathematical formula	$p \propto 1/v$ pv = constant $p_1 v_1 = p_2 v_2 = p_3 v_3 = \cdots = constant$	$v \propto T$ $\frac{v}{T} = constant$ $\frac{v_1}{T_1} = \frac{v_2}{T_2} = \frac{v_3}{T_3} = \cdots = constant$	$p \propto T$ $\frac{p_1}{T_1} = \frac{p_2}{T_2} = \frac{p_3}{T_3} = \cdots = constat$
	T is constant	**p** is constant	**v** is constant