

# **CHM 101:** **Kinetic theory of matter**

(module B)

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## **The Density of a Gas**

One mole of any gas occupies nearly the same volume at a given temperature and pressure, so difference in gas density ( $d = m/V$ ) depend on differences in molar mass.

E.g. 1 mol of  $O_2$  occupies the same volume as 1 mol of  $N_2$ , but since each  $O_2$  molecule has a greater mass than each  $N_2$  molecule,  $O_2$  is denser.

All gasses are miscible when thoroughly mixed, but in the absence of mixing, a less dense gas will lie above a more dense one.

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The ideal gas law can be used to calculate the density of a gas from its molar mass.

$$n = m/M$$

( $n$  = No. of moles;  $m$  = mass;  $M$  = molar mass).

Recall  $PV = nRT$

Therefore,

$$PV = \frac{mRT}{M}$$

Rearrange  $\frac{MP}{RT} = \frac{m}{V} = \text{density}$

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This shows that:

1. *The density of a gas is directly proportional to its molar mass* because a given amount of a heavier gas occupies the same volume as that amount of a lighter gas (Avogadro's law)
2. *The density of a gas is inversely proportional to the temperature.* As the volume of a gas increases with temperature (Charles's law), the same mass occupies more space; thus, the density is lower.

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## **Molar mass**

To determine the molar mass of an unknown gas:

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

$$M = \frac{mRT}{PV}$$

or

$$M = \frac{dRT}{P}$$

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## **The partial Pressure of a gas**

The ideal gas law holds for virtually any gas, whether pure or a mixture, at ordinary conditions for two reasons:

1. Gases mix homogenously (form a solution) in any proportions.
2. Each gas in a mixture behaves as if it were the only gas present (assuming no chemical interactions).

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**Dalton's law of Partial Pressures:** states that "in a mixture of unreacting gases, the total pressure is the sum of the partial pressures of the individual gases".

$$P_{\text{total}} = P_1 + P_2 + P_3 + \dots$$

E.g. suppose a tank of fixed volume contains gas **a** and **b**. Each gas behaves independently,

$$P_a = \frac{n_a RT}{V} \quad \text{and} \quad P_b = \frac{n_b RT}{V}$$

Each gas occupies the same total volume and is at the same temperature, the pressure of a gas depends on its amount,  $n$ .

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Thus, the total pressure is:

$$\begin{aligned} P_{\text{total}} &= P_a + P_b = \frac{n_a RT}{V} + \frac{n_b RT}{V} \\ &= \frac{(n_a + n_b) RT}{V} \\ &= \frac{n_{\text{total}} RT}{V} \end{aligned}$$

$$n_{\text{total}} = n_a + n_b$$

Each component in a mixture contributes a fraction of the total number of moles in the mixture, which is the mole fraction ( $X$ ) of that component.

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For  $n_a$ , the mole fraction is

$$X_a = \frac{n_a}{n_{\text{total}}} = \frac{n_a}{n_a + n_b}$$

Since the total pressure is due to the total number of moles, the partial pressure of gas **a** is the total pressure multiplied by the mole fraction of **a**,

$$P_a = X_a \cdot P_{\text{total}}$$

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### Exercise 8

If 4.58 g of a gas occupies 3.33 L at 27°C and 808 torr, what is the molar mass of the gas?

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### Exercise 9

Calculate the amount of oxygen gas ( $O_2$ ) in a cylinder of 30 L, if the pressure is 20 atm at 30°C.

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### Exercise 10

Calculate the absolute temperature of 0.118 mol of a gas that occupies 10.0 L at 0.933 atm.

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**Exercise 11**

What is the pressure of H<sub>2</sub> if 0.250 mol of H<sub>2</sub> and 0.120 mol of He are placed in a 10.0-L vessel at 27°C.

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**Graham's Law**

**Graham's law** states that "The rate of effusion or diffusion of a gas is inversely proportional to the square root of its mass".

$$\text{Rate of effusion} \propto \frac{1}{\sqrt{m}}$$

$$\text{Rate of diffusion} \propto \frac{1}{\sqrt{m}}$$

**Diffusion** is the passage of a gas through another gas. **Effusion** is the process by which a gas escapes from its container through a tiny hole into an evacuated space.

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Consider two gases with molar masses M<sub>1</sub> and M<sub>2</sub>. The ratio of their rates of diffusion (or effusion) is given by:

$$\frac{r_1}{r_2} = \frac{\sqrt{m_2}}{\sqrt{m_1}}$$

That is, the heavier a molecule of the gas, the more slowly it diffuses (or effuses).

The rate of effusion or diffusion of a gas is directly proportional to the "average" velocity of its molecules.

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**Kinetic Molecular Theory**

The theory is based on 3 postulates (assumptions):

1. **Particle volume**: A gas consists of a large collection of individual particles. The volume of an individual particle is extremely small compared with volume of the container. In essence the model pictures gas particles having mass but no volume.

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2. **Particle motion**: Gas particles are in constant, random, straight-line motion except when they collide with the container walls or with each other.

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3. **Particle collisions**: Collisions are elastic, hence their total kinetic energy is constant. Between collisions, the molecules do not influence each other at all.

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**KINETIC MOLECULAR THEORY ( the model)**

- ✓ Gases are composed of small particles (atoms or molecules).
- ✓ These particles move rapidly in a random, straight line motion. Particles will collide with each other and with the walls of the container.

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- ✓ The bonding forces between particles are extremely weak. It is assumed that particles move around independently.
- ✓ Collisions between particles are elastic, i.e. energy is conserved.

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- ✓ Kinetic energy (energy of movement) can be transferred from one particle to another, but the total kinetic energy will remain constant.
- ✓ The average kinetic energy of the particles increases as the temperature of the gas is increased.

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**Exercise 12**

A sample of calcium carbonate, mass 1.0 g, is heated until it has decomposed completely.

Calculate:

- the mass of carbon dioxide produced
- the volume of carbon dioxide, measured at STP

**SOLN:**

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**Exercise 13**

Copper dispersed in absorbent beds is used to react with oxygen impurities in the ethylene used for producing polyethylene. The beds are regenerated when hot  $H_2$  reduces the metal oxide, forming the pure metal and  $H_2O$ . On a laboratory scale, what volume of  $H_2$  at 765 torr and  $225^\circ C$  is needed to reduce 35.5g of copper (ii) oxide?

**SOLN:**


- To find  $V_{H_2}$ ,  $n_{H_2}$  must first be determined
- Write the equation for d reaction
- Solve for  $n$  (stoichiometry), then solve for  $V$  (gas law)

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**Exercise 14:** What is the density of methane,  $\text{CH}_4$ , at  $20^\circ\text{C}$  and 2.00 atm.

**Exercise 15:** What is the molar mass of a gas whose density at  $40^\circ\text{C}$  and 785 torr is  $1.286 \text{ kg/m}^3$

**Exercise 16:** A collapsed balloon and its load weighs 216 kg. To what volume should it be inflated with H gas in order to launch it from a mountain top at  $-12^\circ\text{C}$  and 628 torr? The density of air under these conditions is  $1.11 \text{ g/L}$ .

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**Exercise 17:** An organic compound containing 55.8% C, 7.03% H, and 37.2% O; was found to have a gas density of  $2.83 \text{ g/L}$  at  $100^\circ\text{C}$  and 740 torr. What is the molecular formula of the compound?

[[Ans: Empirical formula is  $\text{C}_2\text{H}_3\text{O}$ ; molar mass =  $89 \text{ g/mol}$ . Divide  $89 \text{ g/mol}$  by  $43 \text{ g/mol}$  ( $\text{C}_2\text{H}_3\text{O}$  molar mass) = 2. Therefore molecular formula is  $\text{C}_4\text{H}_6\text{O}_2$  ]]

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**Exercise 18:** Chlorine gas is evolved at the anode of a commercial electrolysis cell at the rate of  $3.65 \text{ L/min}$ , at a temperature of  $647^\circ\text{C}$ . On its way to the intake pump it is cooled to  $63^\circ\text{C}$ . Calculate the rate of intake to the pump assuming the pressure has remained constant.

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**Exercise 19:** A spark was passed through a  $50 \text{ cm}^3$  sample of a  $\text{H}_2/\text{O}_2$  mixture in a gas burette at  $18^\circ\text{C}$  and 1.00 atm; the formation of water went to completion. The resulting dry gas had a volume of  $10 \text{ cm}^3$  at  $18^\circ\text{C}$ , 1.00 atm. What was the initial mole fraction of H in the mixture if:

- the residual gas after sparking was  $\text{H}_2$
- the residual gas was  $\text{O}_2$  ?

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**Exercise 20:** A rigid plastic container holds  $35.0 \text{ g}$  of ethylene gas ( $\text{C}_2\text{H}_4$ ) at a pressure of 793 torr. What is the pressure if  $5.0 \text{ g}$  of ethylene is removed at constant temperature?

**Exercise 21:** A scale model of a blimp rises when it is filled with helium to a volume of  $55.0 \text{ dm}^3$ . When  $1.10 \text{ mol}$  of He is added to the blimp, the volume is  $26.2 \text{ dm}^3$ . How many more grams of He must be added to make it rise? Assume constant T and P.

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