

Module 3: Representing Amounts of Substances

Amounts of Gases; The Ideal Gas Law

Fundamentals of Chemistry Open Course

1. Explain the significance of Avogadro's number and why the value $6.022 \times 10^{23} \text{ mol}^{-1}$ is a convenient definition of the mole.
2. Use average atomic masses to calculate the molar mass of a substance with given chemical formula.
3. Apply molar mass to determine amount of substance from mass and *vice versa*.
4. Define mass density and molar volume; apply them in calculations.
5. Visualize liquid solutions at the submicroscopic level; identify the components of a solution.
6. Define concentration and recognize common units of concentration.
7. Define molarity and apply it to calculate amount of solute from volume of a solution and *vice versa*.
8. Recognize quantities in the ideal gas law and their associated units.
9. Apply the ideal gas law to calculate the amount of a gas from pressure, volume, and temperature.

- **Gases** are a dispersed phase in which atoms or molecules are very far apart from one another and in constant motion.
- Particles of a gas are uniformly distributed over space; gases expand to fill the volume of their containers.
- Particles striking the walls of the container exert a force; this force divided by the area over which it acts is called **pressure**.
- Units of pressure include...
 - Force per area units: $\text{N/m}^2 = \text{Pa}$, $\text{lb/in}^2 = \text{psi}$, bar, etc.
 - Height of a column of fluid (**hydrostatic pressure**): mmHg , inH_2O , ftH_2O
 - Atmospheric pressure: atm

Ideal Gases

- The **ideal gas model** is an extremely important submicroscopic model of gas behavior.
- The model assumes no interactions between gas particles and collisions (between particles or between a particle and a wall) that conserve total kinetic energy.
- Gas particles are not all moving at the same speed; particles display a distribution of speeds.
- Play around with the simulation linked below to get a feel for this model.

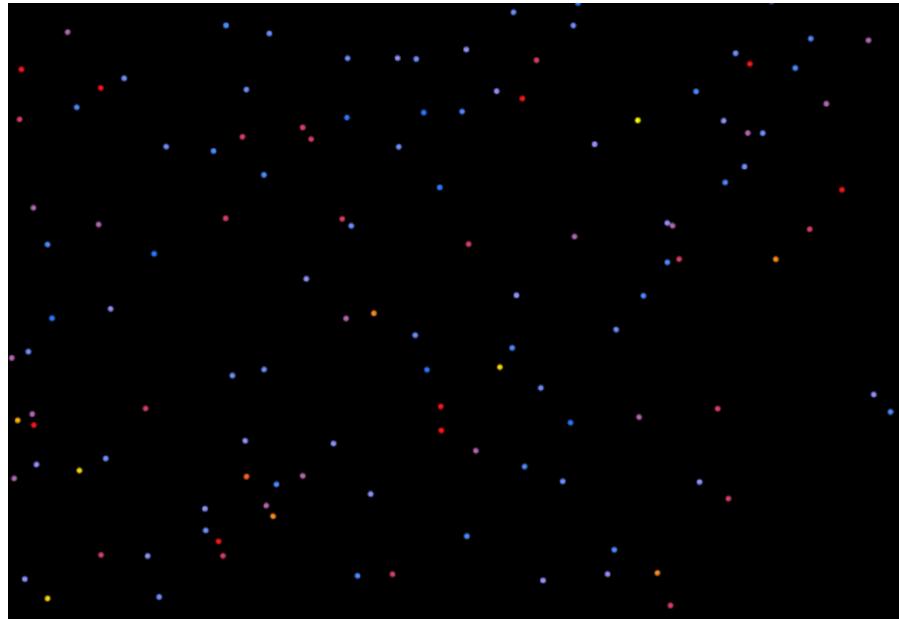


Figure. Submicroscopic model of an ideal gas.
See falstad.com/gas for an interactive simulation.

The Ideal Gas Law

- From the ideal gas model comes an equation that relates amount in moles n , pressure P , volume V , and absolute temperature T : the **ideal gas law** or **ideal gas equation of state**.

$$PV = nRT$$

- The **gas constant R** has a value of $0.08206 \text{ L atm K}^{-1} \text{ mol}^{-1}$.
- The key idea for us here is that if P , V , and T are known, the amount of gas can be calculated using the ideal gas law!

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

Example. How many moles of O_2 gas are present in a 2.00-L cylinder at 293 K and 1.00 atm pressure?

Proportional Reasoning with Ideal Gases

$$PV = nRT$$

- From the ideal gas law, it follows that...
 - At constant T and P , volume is proportional to number of moles. $V_1/n_1 = V_2/n_2$
 - At constant T and V , pressure is proportional to number of moles. $P_1/n_1 = P_2/n_2$
- If we know the pressure P_1 corresponding to a given amount in moles n_1 at constant temperature and volume, we can find the pressure P_2 at a different amount in moles n_2 .
- If we know the volume V_1 corresponding to a given amount in moles n_1 at constant temperature and pressure, we can find the volume V_2 at a different amount in moles n_2 .

Example. A sample of 0.352 mol H₂ has a pressure of 1.86 atm. At what number of moles would this gas display a pressure of 3.12 atm, assuming no change in temperature or volume?

Lingering Questions

- What other relations between gas variables can we infer from the ideal gas law? How might we use these equations?
- Does the ideal gas law apply to a mixture of two or more gases? Why or why not?
- How can we use the ideal gas law to plan and carry out reactions involving gases?
- What happens when a gas deviates from ideal behavior? How do we model real gases?