CHEMISTRY

an atoms-focused approach

Gilbert
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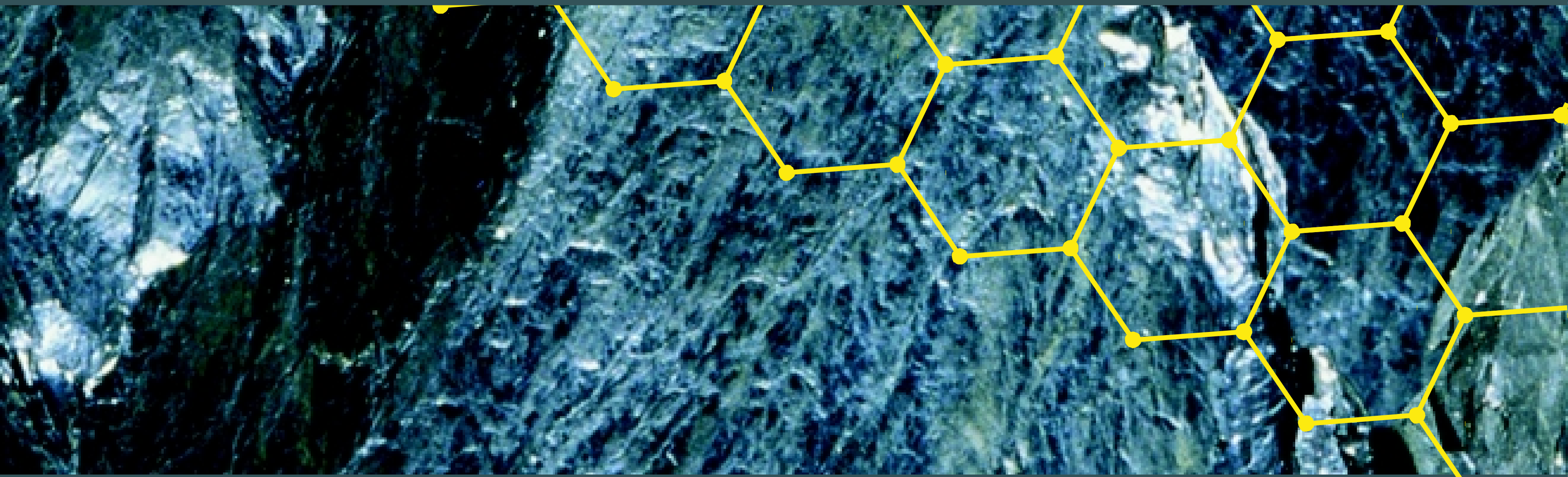


Chapter 10

Properties of Gases

The Air We Breathe

Chapter Outline

10.1 The Properties of Gases

10.2 Effusion and the Kinetic Molecular Theory of Gases

10.3 Atmospheric Pressure

10.4 The Gas Laws

10.5 The Combined Gas Law

10.6 Ideal Gases and the Ideal Gas Law

10.7 Densities of Gases

10.8 Gases in Chemical Reactions

10.9 Mixtures of Gases

10.10 Solubility of Gases and Henry's Law

10.11 Gas Diffusion: Molecules Moving Rapidly

10.12 Real Gases

Properties of a Gas

- Neither definite shape nor definite volume
  - Uniformly fills any container
  - Exerts pressure on surroundings
  - Volume changes with temperature and pressure
- Mixes completely with other gases
- Much less dense than solids, liquids

Parameters Affecting Gases

- Pressure (*P*)
- Volume (*V*)
- Temperature (*T*)
- Number of Moles (*n*)

Chapter Outline

Kinetic Molecular Theory (KMT)

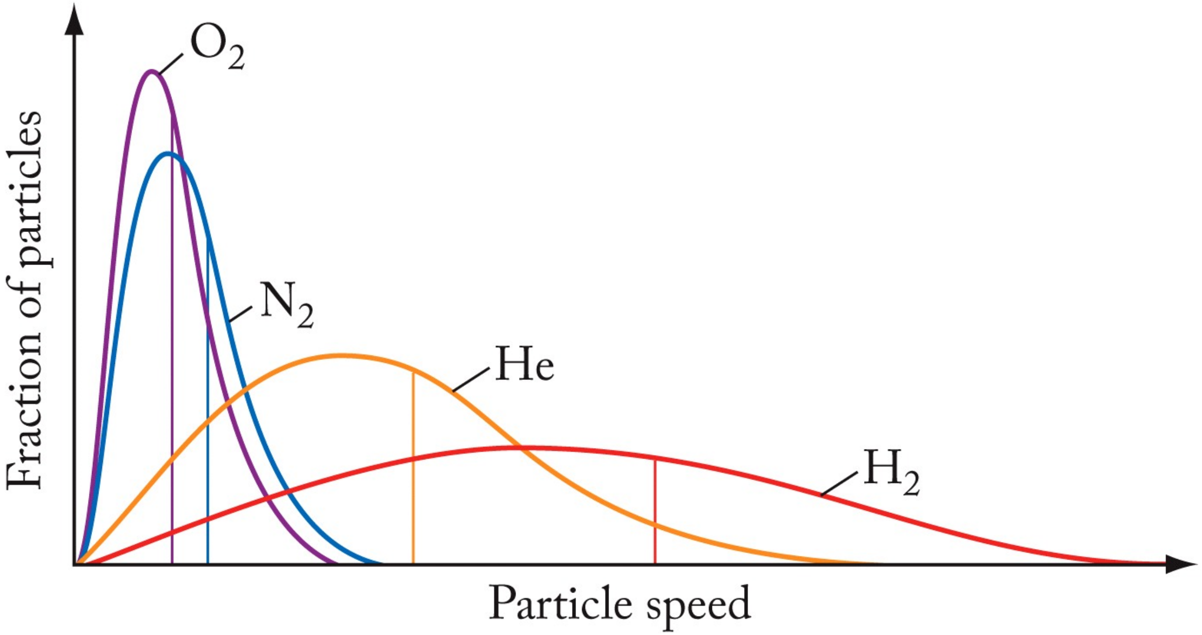
Assumes that gas molecules:

1. Have tiny volumes compared with their container's volume
2. Don't interact with other gas molecules
3. Move randomly and constantly
4. Engage in elastic collisions with walls of container and other gas molecules
5. Have average kinetic energy that is proportional to absolute temperature

Kinetic Molecular Theory (cont.)

- Average Kinetic Energy: $KE_{avg} = \frac{1}{2}\, m u^2_{rms}$
  - u_{rms} = the root-mean-squared speed of the molecules;

    m = molecular mass.



Effusion

- Relative Rates of Effusion:

$$\frac{(\text{Rate})_{\text{gas 1}}}{(\text{Rate})_{\text{gas 2}}} = \sqrt{\frac{M_2}{M_1}}$$

where M is the molar mass

Chapter Outline

The Gas Phase

- The Atmosphere:
  - Layer of gases 50 km thick
  - Composition is fairly consistent
  - Properties vary with location
    - » Pressure, density

TABLE 10.1 Composition of Dry Air[a]

Compound	% (by volume)
Nitrogen	78.08
Oxygen	20.95
Argon	0.934
Carbon dioxide	0.0395[b]
Neon	0.0018
Helium	0.00052
Methane	0.00018
Krypton	0.00011

[a]Includes major and minor gases (with concentrations >1 ppm by volume).
[b]Value as of January 2013. Atmospheric CO_2 is increasing by about 2 ppm each year.

Pressure

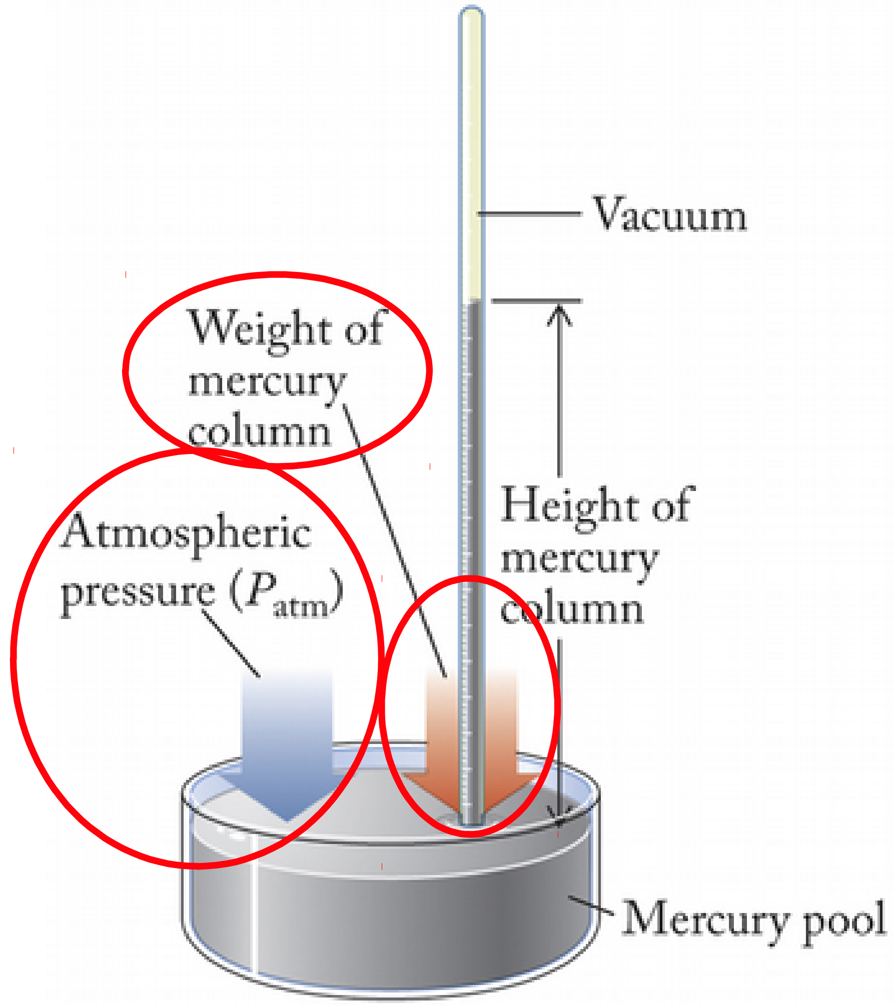
- Pressure:
  - Force/unit area ($P = F/A$)
  - Atmospheric pressure = pressure exerted due to gravity acting on air above Earth's surface
- Units of Pressure:
  - SI units: newton/meter^2 = 1 pascal (Pa)
  - 1 standard atmosphere (1 atm) = 101,325 Pa
  - 1 atm = 760 mmHg = 760 torr

Measurement of Pressure

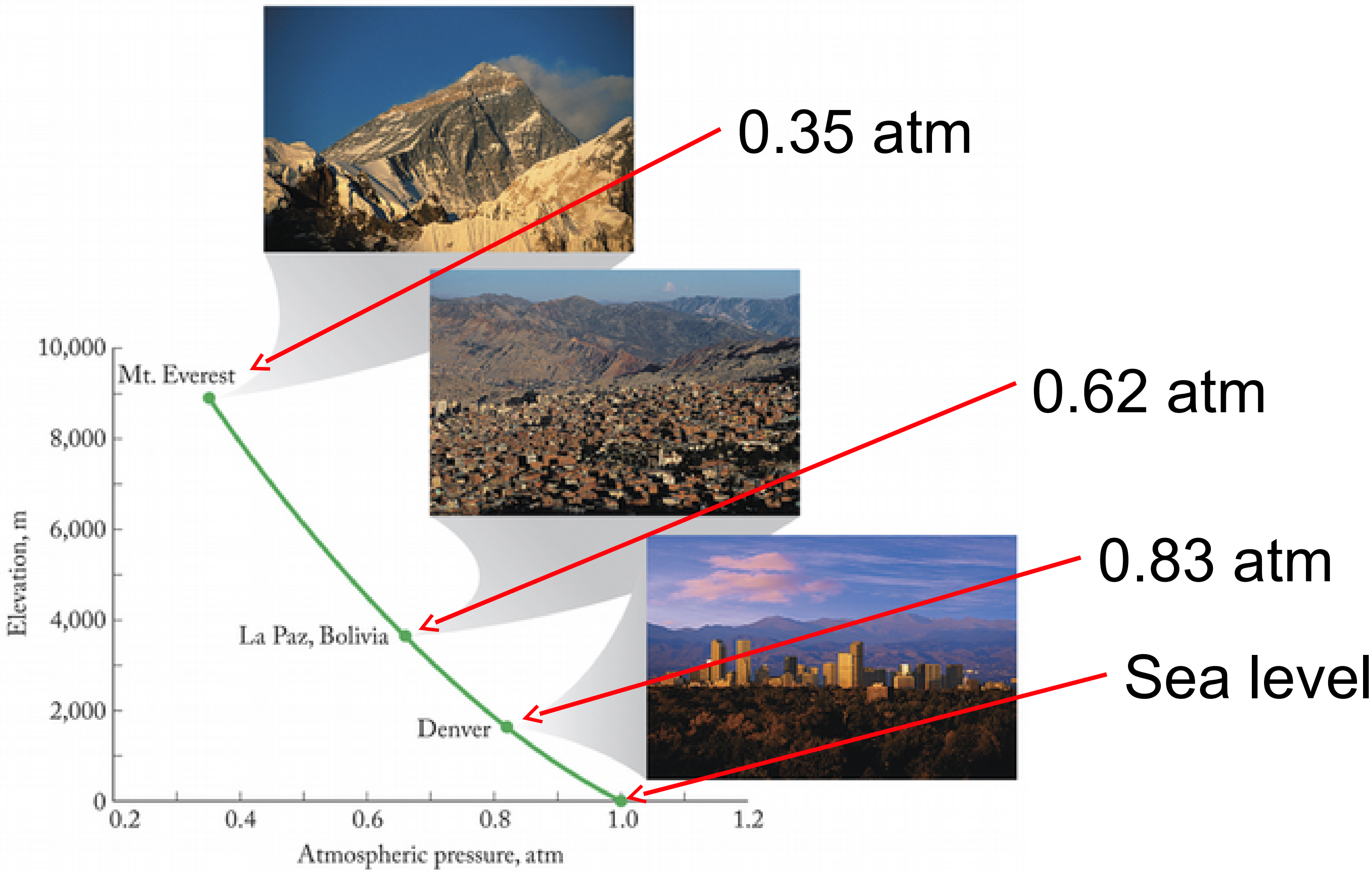
Barometer: measures atmospheric pressure

Height of Hg column based on balance of forces:

- gravity (pulls Hg down).
- atmospheric pressure (pushes Hg up into evacuated tube)



Elevation and Atmospheric Pressure

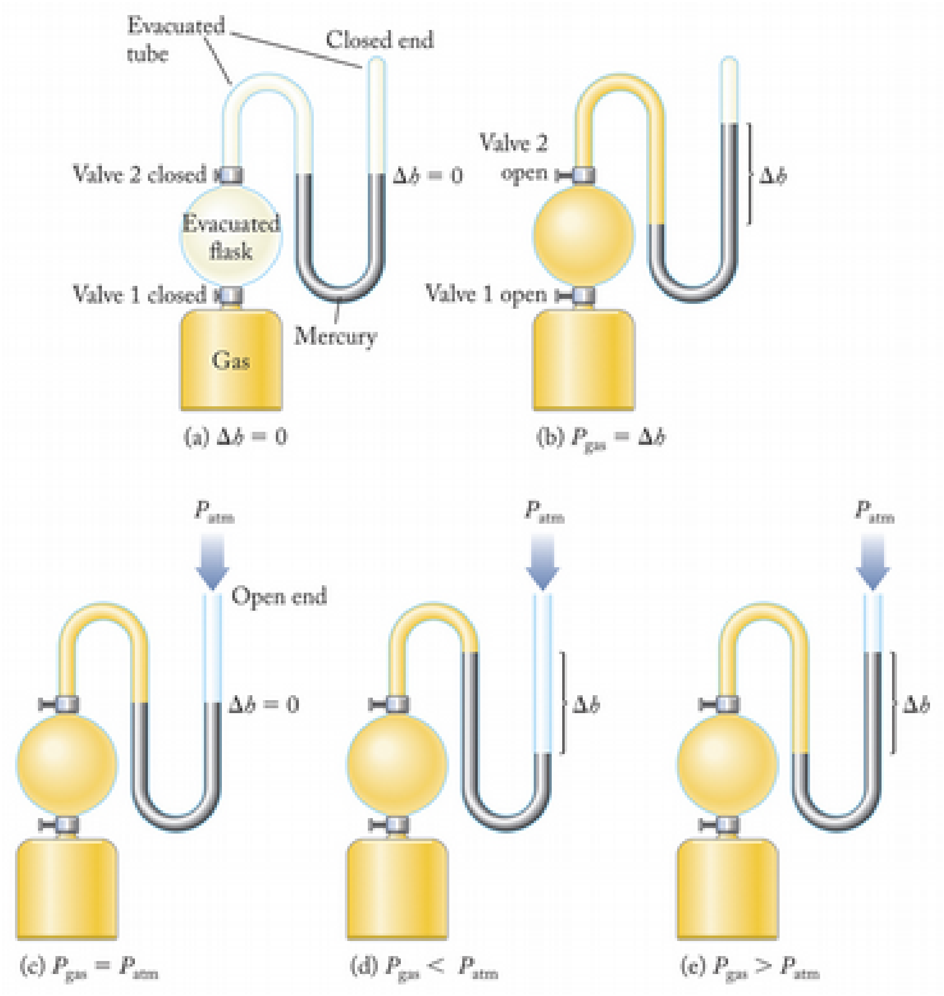


Relationship between Pressure Units

TABLE 10.2 Units for Expressing Pressure

Unit	Value
Standard atmosphere (atm)	1 atm
Pascal (Pa)	1 atm = 1.01325×10^5 Pa
Kilopascal (kPa)	1 atm = 101.325 kPa
Millimeter of mercury (mmHg)	1 atm = 760 mmHg
Torr	1 atm = 760 torr
Bar	1 atm = 1.01325 bar
Millibar (mbar or mb)	1 atm = 1013.25 mbar
Pounds per square inch (psi)	1 atm = 14.7 psi
Inches of mercury	1 atm = 29.92 inches of Hg

Measuring Pressure: Manometer



Open systems:

- Δh is negative if $P_{gas} < P_{atm}$ (d)
- Δh is positive if $P_{gas} > P_{atm}$ (e)

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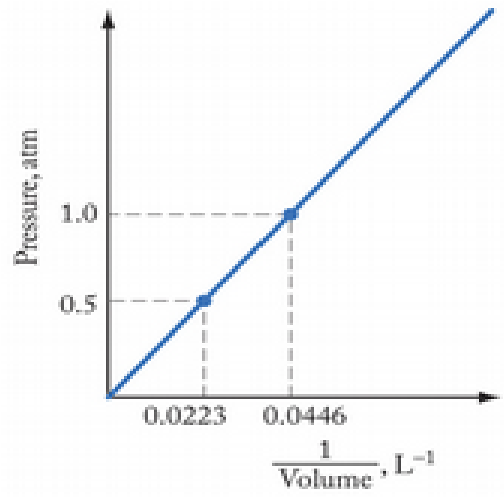
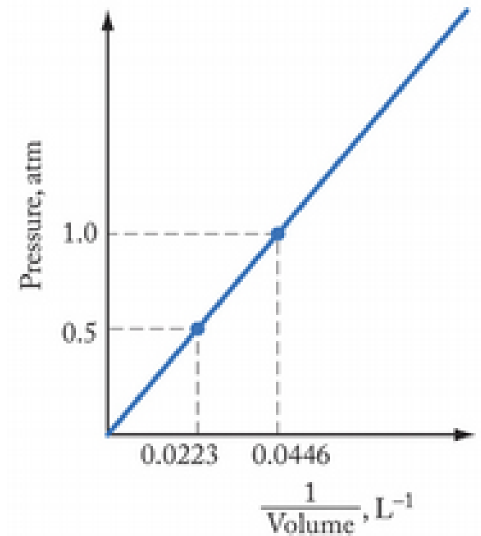
Boyle's Law

- Gases are compressible
  - Pressure ↑ as Volume ↓
- Boyle's Law:
  - $P \propto 1/V$ (T and n fixed)
  - or, $P \times V$ = constant
  - or, $P_1V_1 = P_2V_2$
- Decreasing volume increases number of collisions/area; P↑ (KMT Postulates #3 & 4)





Practice: Boyle's Law

A balloon is filled with carbon dioxide to a pressure of 1.85 atm and has a volume of 1.54 L. If temperature remains constant, what is the final volume when the pressure is increased to 2.50 atm?

- Collect and Organize:
- Analyze:
- Solve:
- Think about It:

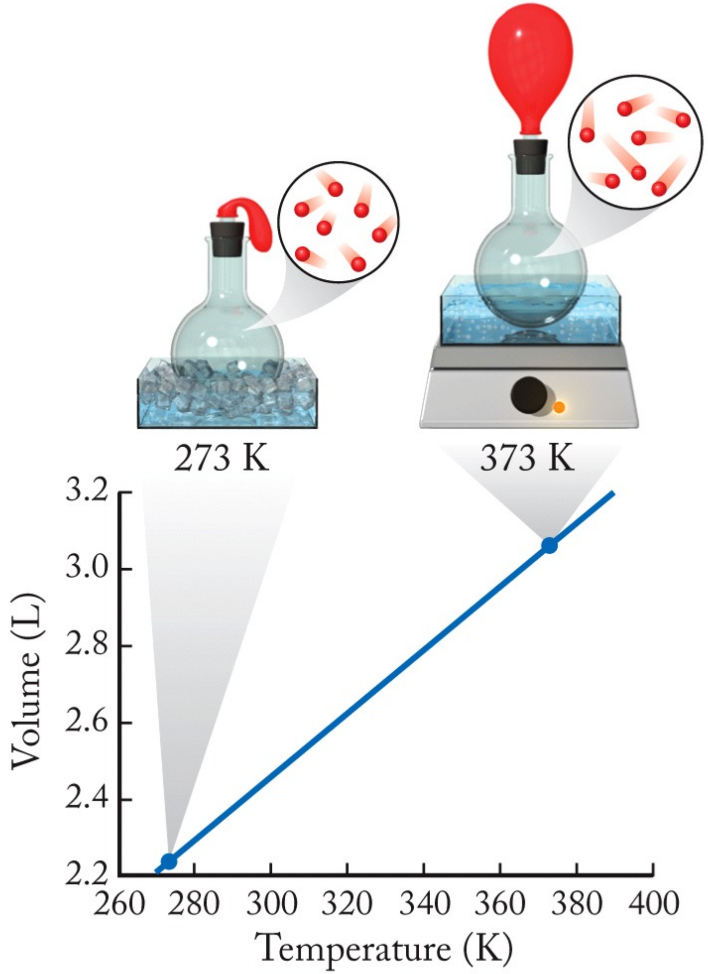
Charles's Law

- Charles's Law:
  - $V \propto T$ (P, n constant)

    or, $$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

    Volume of a gas extrapolates to zero at absolute zero (0 K = −273°C).

    Kinetic energy ↑ as T ↑; force of collisions increases and gas expands to maintain constant *P* (KMT Post. #3, 4 & 5).

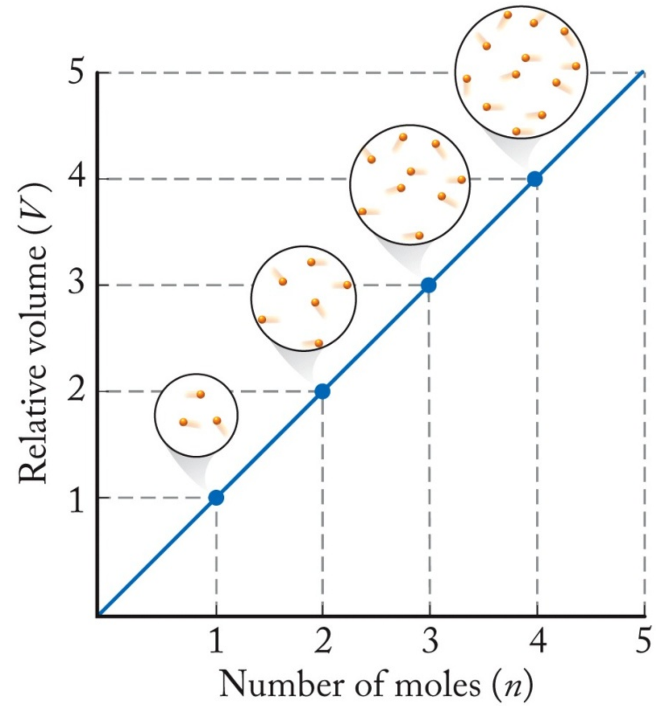


Avogadro's Law

- Volume is directly proportional to the number of moles of gas, $V \propto n$ (T, P constant)

$$or, \quad \frac{V_1}{n_1} = \frac{V_2}{n_2}$$

$$or, \quad \frac{V_1}{n_1} = \frac{V_2}{n_2}$$



Increasing *n* increases the number of collisions, gas expands to keep pressure constant (KMT Post. #3 & 4).

Amonton's Law

- $P \propto T$ (n, V constant)

$$\frac{P}{T} = \text{constant}$$

$$\frac{P}{T} = \text{constant}$$

$$\frac{P}{T} = \text{constant}$$

Increasing T will increase force of collisions if volume is kept constant; P will increase (KMT Post. #3, 4 & 5).

Chapter Outline

Combined Gas Law

- Boyle's Law: $P \times V$ = constant
- Charles's Law: V/T = constant
- Avogadro's Law: V/n = constant
- Combining the gas laws: $\dfrac{P \cdot V}{n \cdot T} = \text{constant}$
- If n is constant, then PV/T = constant, and

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$$

Practice: Combined Gas Law

A sample of oxygen gas is at 0.500 atm and occupies a volume of 10.0 L at 0°C. What is the pressure of the gas if it is at 15.0 L at 25°C?

- Collect and Organize:
- Analyze:
- Solve:
- Think about It:

Chapter Outline

Ideal Gas Law

Combined Gas Law: $$\frac{P \cdot V}{n \cdot T} = \text{constant} = R$$

This rearranges to: $PV = nRT$

R = universal gas constant = 0.08206 L·atm K^{-1} mol^{-1}

P = pressure (in atm)

V = volume (in liters)

n = moles

T = temperature (in kelvin)

Universal Gas Constant

- Value of universal gas constant depends on the units
- Using SI units of *V* and *P*,

$$R = 0.08206\ \frac{\mathrm{L \cdot atm}}{\mathrm{mol \cdot K}}$$

- Using 1 atm = 101.325 kPa to convert to atm

$$R = 0.08206\ \frac{\mathrm{L \cdot atm}}{\mathrm{mol \cdot K}}$$

Practice: Ideal Gas Law

Calculate the moles of gas contained in a 4.0 L container at STP.

- Collect and Organize:
- Analyze:
- Solve:
- Think about It:

Practice: Ideal Gas Law

Calculate the pressure of 4.0 mol of methane gas in a 12.3 L container at 25°C.

- Collect and Organize:
- Analyze:
- Solve:
- Think about It:

Practice: Ideal Gas Law

An experiment shows that a 0.281 g sample of an unknown gas occupies 127 mL at 98°C and 754 torr pressure. Calculate the molar mass of the gas. (Hint: $\mathcal{M} = g/n$)

- Collect and Organize:
- Analyze:
- Solve:
- Think about It:

Reference Points for Gases

- Standard Temperature and Pressure (STP):
  - P = 1 atmosphere; T = 0°C (273 K)
- Molar Volume:
  - For 1 mol of an ideal gas at STP (calculated from the ideal gas law):

    V = 22.4 L

Chapter Outline

Densities of Gases

- Can be calculated from molar mass ($\mathcal{M}$) and molar volume (*V*/*n*).
- From Ideal Gas Law:
  - $PV = nRT \rightarrow$ $$d = \frac{m}{V} = \frac{P\mathcal{M}}{RT}$$
  - Density: $$d = \frac{m}{V} = \frac{P\mathcal{M}}{RT}$$
  - When *P* in atm, *T* in kelvin, *d* = g/L

Buoyancy: Densities of Gases

Buoyancy depends on differences in gas densities.

Depends on:

1. Molar Masses

   He(*g*) = 0.169 g/L*

   N_2(*g*) = 1.19 g/L*

2. Temperature

   Charles's Law:
   density ↓ as Temp.↑

* At 15°C and 1 atm

Practice: Densities of Gases

When HNO_3(*aq*) and $NaHCO_3$(*aq*) are mixed together, a reaction takes place in which a gas is one of the products. The gas has a *d* = 1.83 g/L at 1.00 atm and 23°C. What is the molar mass and identity of the gas?

- Collect and Organize:
- Analyze:
- Solve:
- Think about It:

Chapter Outline

Stoichiometric Calculations Using Gases

- Stoichiometric Calculations:
  - Depend on mole/mole ratios of reactants and/or products
  - Moles of gas can be calculated from ideal gas law if *P*, *V*, and *T* are known

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

Practice: Gas Stoichiometry

Automobile air bags inflate during a crash or sudden stop by the rapid generation of $N_2(g)$ from sodium azide:

$$2NaN_3(s) \rightarrow 2Na(s) + 3N_2(g)$$

How many grams of sodium azide are needed to produce sufficient $N_2(g)$ to fill a 48 × 48 × 25 cm bag to a pressure of 1.20 atm at 15°C?

- Collect and Organize:
- Analyze:
- Solve:
- Think about It:

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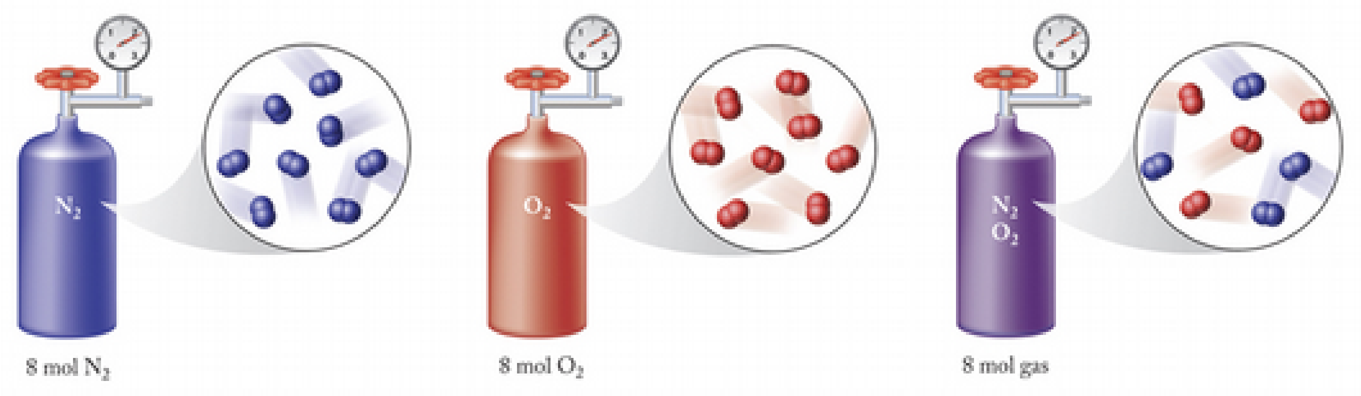
Dalton's Law of Partial Pressures

- For a mixture of gases in a container:
  - $P_{total} = P_1 + P_2 + P_3 + \ldots$



Total pressure depends only on total number moles of gas, not on their identities (KMT Post. #2).

Mole Fraction & Partial Pressure

- Mole Fraction:
  - Ratio of the # of moles of a given component in a mixture to the total # of moles in a mixture:

$$x_1 = \frac{n_1}{n_{\text{total}}} = \frac{n_1}{n_1 + n_2 + n_3 + \ldots}$$

- Mole Fraction in Terms of Pressure:
  - When *V* and *T* are constant, $P \propto n$

Mole Fraction & Partial Pressure (cont'd)

Since $P \propto n$

$$x_1 = \frac{P_1}{P_{\text{total}}}$$

And:

$$x_1 = \frac{P_1}{P_{\text{total}}}$$

Then…

$$x_1 = \frac{P_1}{P_{\text{total}}}$$

Practice: Mole Fraction

At 25°C, a 1.0 L flask contains 0.030 mol of oxygen, 150.0 mg of nitrogen, and 2.6×10^{21} molecules of carbon dioxide. Calculate the partial pressure and mole fraction of each gas.

- Collect and Organize:
- Analyze:
- Solve:
- Think about It:

Collecting a Gas over Water

$$2KClO_3(s) \rightarrow 2KCl(s) + 3O_2(g)$$

Gases collected:

$O_2(g)$ and $H_2O(g)$

$$P_{\text{total}} = P_{O_2} + P_{H_2O}$$

$$P_{\text{total}} = P_{O_2} + P_{H_2O}$$

Practice: Partial Pressure of Water

A sample of $KClO_3$ is heated and decomposes to produce O_2 gas. The gas is collected by water displacement at 25°C. The total volume of the collected gas is 329 mL at a pressure of 744 torr. How many moles of oxygen are formed?

- Collect and Organize:
- Analyze:
- Solve:
- Think about It:

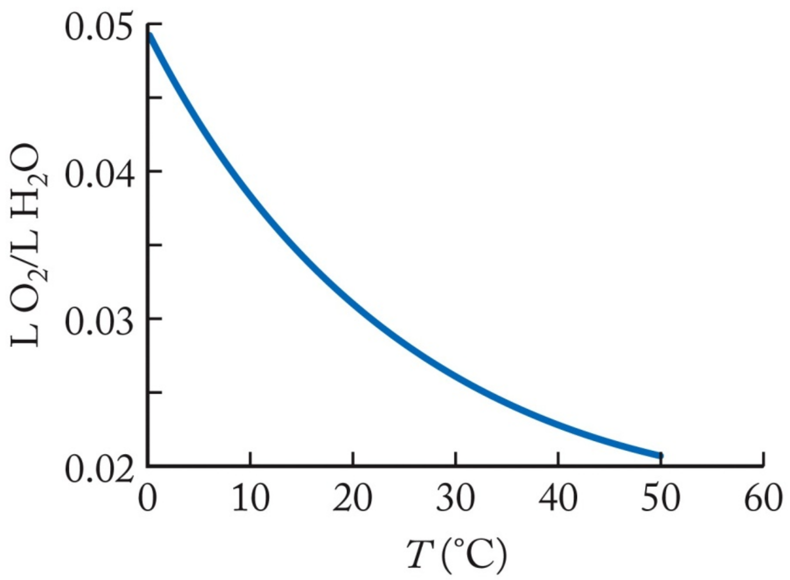
Chapter Outline

Solubility of Gases

- Solubility of gases depends on *T* and *P*

Solubility ↑ as Pressure ↑

Solubility ↓ as Temperature ↑



Henry's Law

- Henry's Law:
  - *The higher the partial pressure of the gas above a liquid, the more soluble*
  - $C_{gas} \propto P_{gas}$
  - $C_{gas} = k_H P_{gas}$

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Diffusion and Effusion

- Graham's Law:
  - Rate of Effusion and Diffusion $\propto \frac{(\text{Distance})_{\text{gas 1}}}{(\text{Distance})_{\text{gas 2}}} = \sqrt{\frac{M_1}{M_2}}$
  - Relative Rates of Effusion: $$\frac{(\text{Distance})_{\text{gas 1}}}{(\text{Distance})_{\text{gas 2}}} = \sqrt{\frac{M_1}{M_2}}$$
  - Diffusion (Distance): $$\frac{(\text{Distance})_{\text{gas 1}}}{(\text{Distance})_{\text{gas 2}}} = \sqrt{\frac{M_1}{M_2}}$$

Practice: Graham's Law

List the following gases, which are at the same temperature, in the order of increasing rates of diffusion: He, Kr, NO, O_2

- Collect and Organize:
- Analyze:
- Solve:
- Think about It:

Practice: Graham's Law

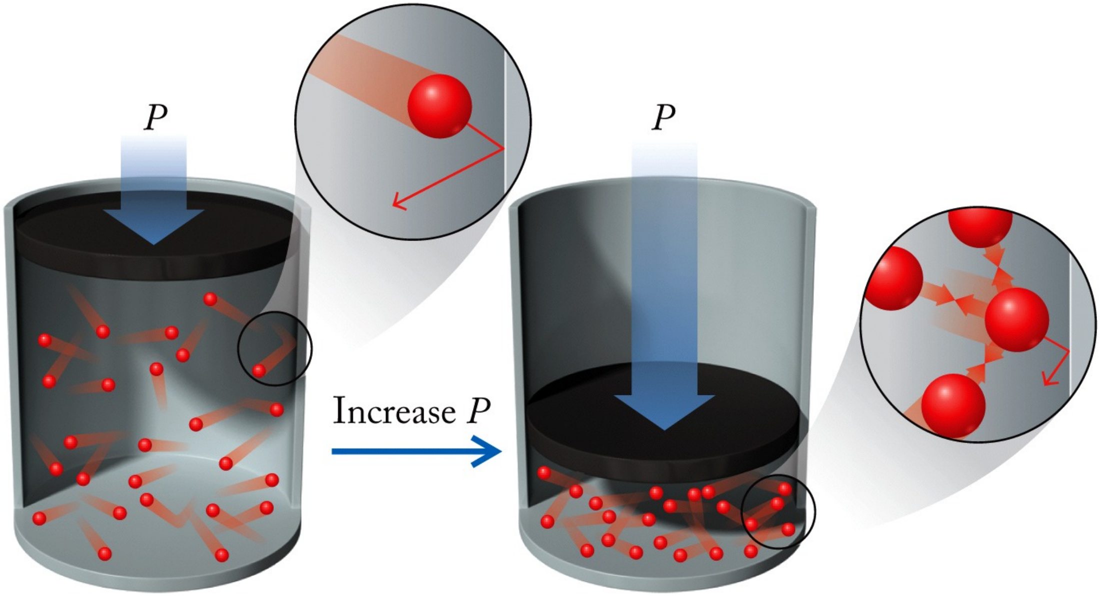
Calculate the molar mass of a gas if equal volumes of oxygen gas and the unknown gas take 3.25 and 4.60 min, respectively, to effuse through a small hole at constant pressure and temperature.

- Collect and Organize:
- Analyze:
- Solve:
- Think about It:

Chapter Outline

Real Gases

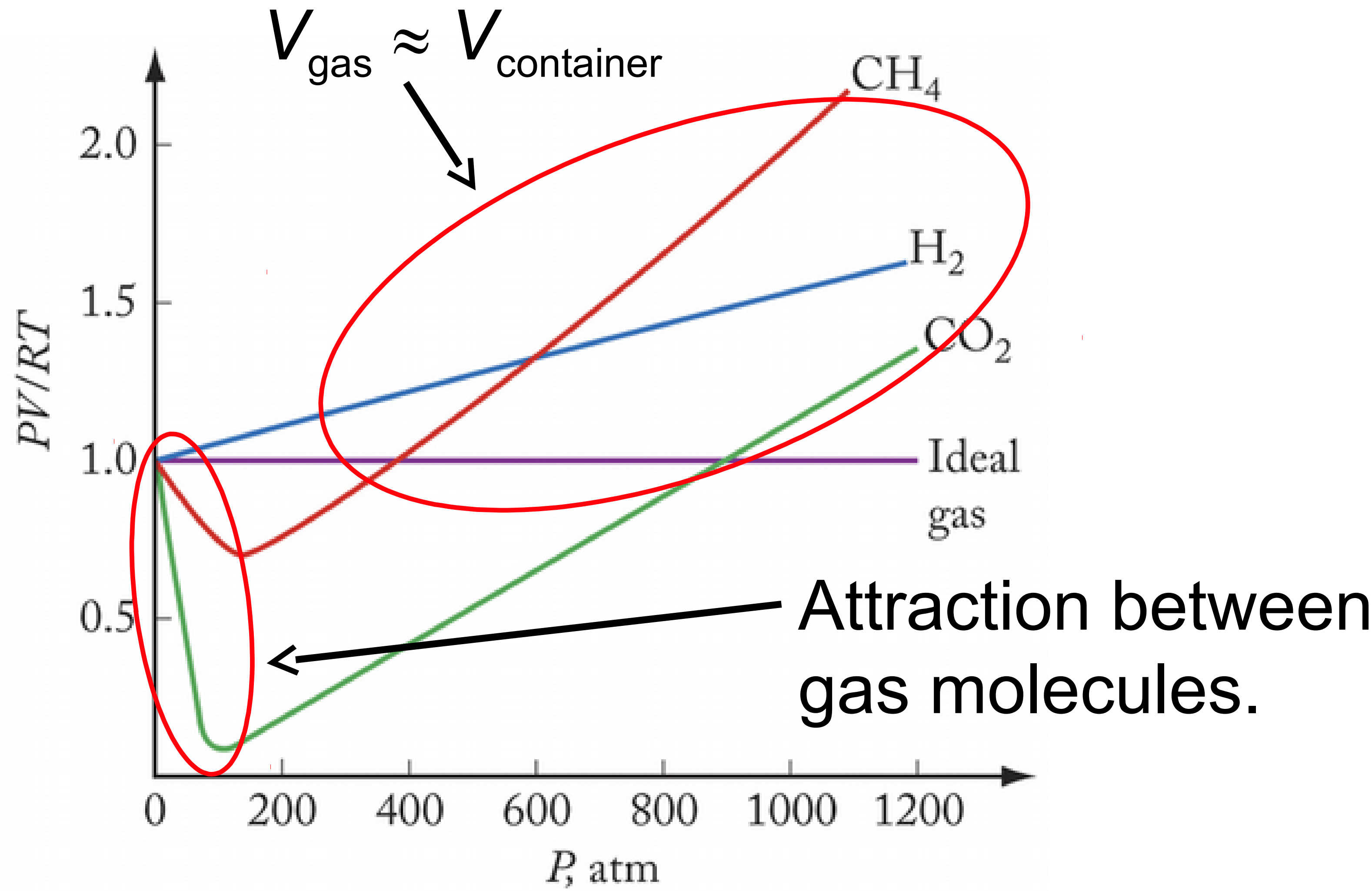
Ideal gas behavior must be corrected when at high pressure (smaller volume) and low temperature (attractive forces become important).



Ideal vs. Real Gases

- Assumptions of Kinetic Molecular Theory:
  - #1: V_{gas} is negligible compared to $V_{container}$.
  - #5: Gas molecules act independently (i.e., don't interact with each other).
- Valid at STP, but not at higher pressures:
  - Volume occupied by gas molecules is not negligible.
  - Attractive forces between gas molecules are significant.

Deviations from Ideal Behavior



Real Gases

- **Corrections to Ideal Gas Law:**
  - **van der Waals Equation**

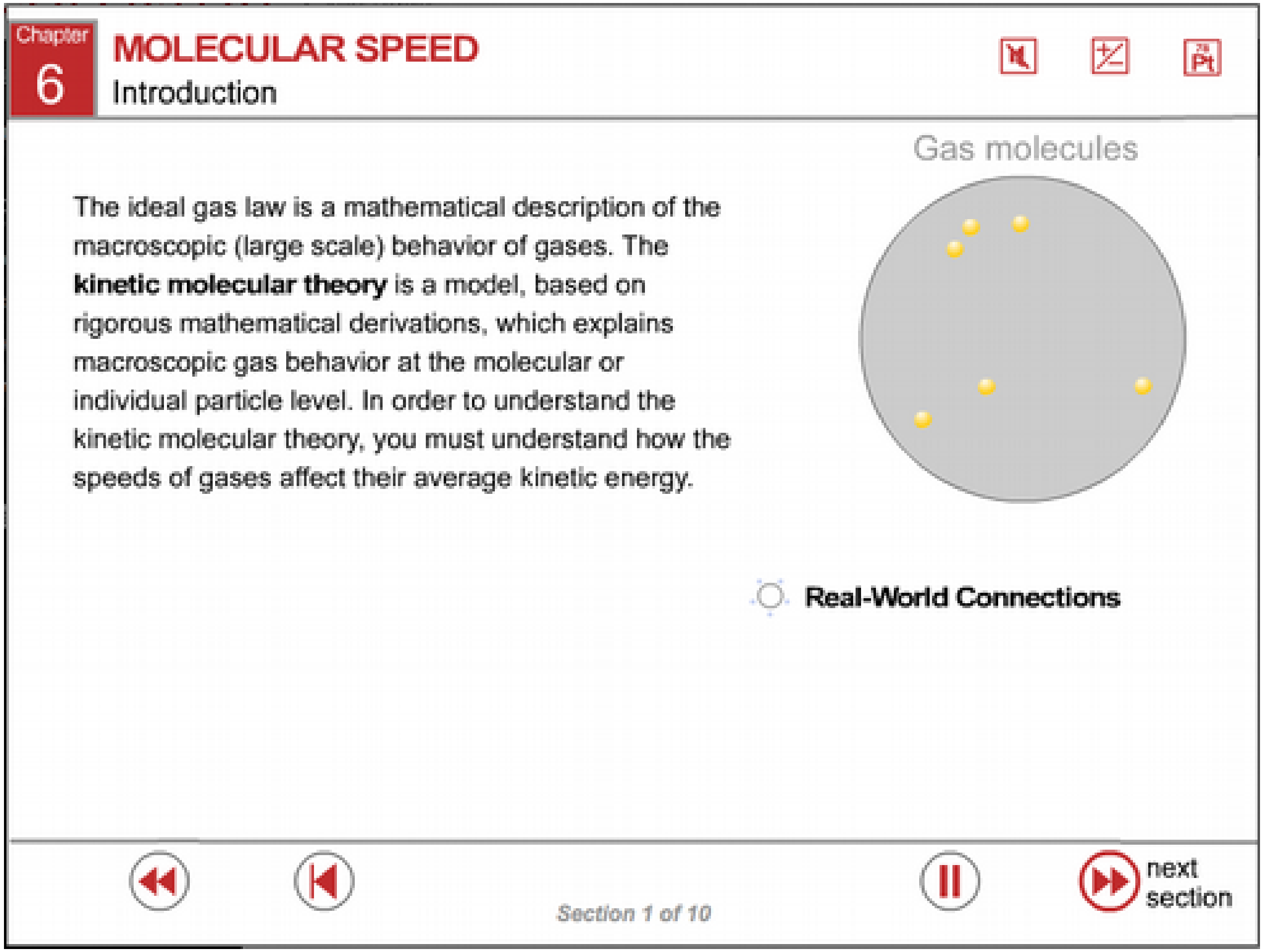
$$\left(P + \frac{n^2 a}{V^2}\right)(V - nb) = nRT$$

$$\underbrace{\left(P + \frac{n^2 a}{V^2}\right)}_{P_{ideal}} \quad \underbrace{(V - nb)}_{V_{ideal}}$$

corrected pressure → P_{ideal}

corrected volume → V_{ideal}

ChemTours: Chapter 10



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This concludes the Lecture PowerPoint presentation for Chapter 10

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