

University Chemistry I
Chemistry 120, Section A
Chapter 5 — Gases

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5-1

A Gas

- Uniformly fills any container
- Mixes completely with any other gas
- Exerts pressure on its surroundings

5-2

Pressure

- is equal to force/unit area
- SI units = Newton/meter² = 1 Pascal (Pa)
- 1 standard atmosphere (atm) = 101,325 Pa
- 1 atm = 760 mm Hg = 760 torr

5-3

Gas Laws: Boyle's Law*

Boyle's Law- At constant temperature, volume is inversely proportional to pressure.

$$V \propto \frac{1}{P} \quad \text{or} \quad P_1V_1 = P_2V_2$$

*Holds *precisely* only at very low pressures.

5-4

Ideal Gas

- A gas that strictly obeys Boyle's Law is called an ideal gas.
- (More later)

5-5

Gas Laws: Gay-Lussac's Law

Gay-Lussac's Law- At constant V

$$P \propto T \quad \text{or} \quad \frac{P_1}{T_1} = \frac{P_2}{T_2}$$

5-6

Measurement: Temperature

Must be in Kelvins

$$^{\circ}\text{C} + 273.15 = \text{Kelvins}$$

5-7

Gas Laws: Charles' Law

Charles' Law- At constant pressure the volume of a gas is directly proportional to temperature, and extrapolates to zero at zero Kelvin.

$$V \propto T \quad \text{or} \quad \frac{V_1}{T_1} = \frac{V_2}{T_2}$$

$$V = \kappa T$$

5-8

Combined Gas Law

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

Where P= pressure, V=volume, and T=temp (K)

5-9

Practice Problems: Combined Gas Law

What temperature is required to cause the pressure of a (steel) cylinder of gas to increase from 350 to 500 mm Hg? The initial temperature was 298 K.

5-10

Practice Problems: Combined Gas Law

Suppose you have 856 mL of a gas. A weather front comes through, and the barometric pressure changes from 780 mm Hg to 720 mm Hg. Along with this, the temperature changes from 86 °F (30. °C) to 72 °F (22 °C). What is the new volume of your gas?

5-11

Avogadro's Law

For a gas at constant temperature and pressure, the volume is directly proportional to the number of moles of gas (at low pressures).

$$V = a n$$

a = proportionality constant

V = volume of the gas

n = number of moles of gas

1 mol occupies 22.42 L at *STP*

5-12

Standard Temperature and Pressure (STP)

$P = 1$ atmosphere

$T = 0$ °C

The molar volume of an ideal gas is 22.42 liters at STP.

1 mol occupies 22.42 L at *STP*

5-13

Ideal Gas Law

$$PV = nRT$$

The molar volume of an ideal gas is 22.42 liters at STP.

From this, can calculate *universal gas constant*, R .

5-14

Ideal Gas Law

$$PV = nRT$$

$$R = \text{proportionality constant} \\ = 0.08206 \text{ L atm K}^{-1} \text{ mol}^{-1} \quad R = \frac{0.08206 \text{ L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}$$

P = pressure in atm

V = volume in liters

n = moles

T = temperature in Kelvins

Holds closely at $P < 1$ atm

5-15

Practice Problems - Ideal Gas Law

1. What volume will 52.5 g of CH_4 occupy at STP?
2. You heat 1.437 g NH_3 in a stoppered 250 mL flask until it explodes (425 °C). What was the pressure inside the flask immediately prior to the explosion?

5-16

Molar Mass of a Gas — Exp 07

- In a 24.5 °C laboratory with a barometric pressure of 782 mm Hg, a 250-mL Ehrlenmeyer flask filled with an unknown gas was found to contain 0.448 g gas. The same flask filled with water was found to have a mass of 265.545 g. The density of water at 24.5 °C is 0.9971 g/mL. What is the molar mass of the gas?
- What is the density of this gas?

5-17

Dalton's Law of Partial Pressures

For a mixture of gases in a container, the total pressure of a gas mixture equals the sum of the partial pressures of each gas in the mixture.

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

5-18

Practice Problems: Dalton's Law of Partial Pressures

What is the total pressure of a mixture of He and N₂ if their partial pressures are 240 and 850 mm Hg?

A gas mixture with 80% He and 20% O₂ has a total pressure of 800 mm Hg. What is the partial pressure of O₂?

If the barometric pressure changes from 780 mm Hg to 720 mm Hg, how much does the pressure of N₂, which is 78% of the atmosphere, change?

5-19

Molar Volume for Several Gases at STP

Gas	Molar Volume L
Hydrogen (H ₂)	22.433
Helium (He)	22.434
Nitrogen (N ₂)	22.402
Oxygen (O ₂)	22.397
Carbon dioxide (CO ₂)	22.260
Ammonia (NH ₃)	22.079

5-20

Kinetic Molecular Theory

1. All particles are in continuous, random motion.
2. Particles move in a straight line and change directions only when a collision occurs.
3. Collisions are elastic.
4. Volume of individual particles is \approx zero.
5. Collisions of particles with container walls cause pressure exerted by gas.
6. Particles exert no forces on each other.
7. Average kinetic energy \propto Kelvin temperature of a gas.

5-21

The Meaning of Temperature

$$KE_{\text{avg}} = \frac{3}{2} RT$$

Kelvin temperature is an index of the random motions of gas particles (higher T means greater motion).

5-22

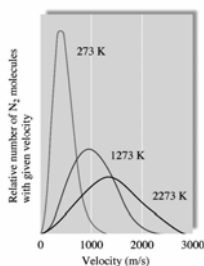
$$u_{\text{rms}}$$

Fig. 5.21 Zumdahl and Zumdahl, 6th Ed.

- Not all molecules travel at the same speed
- Averaged speed denoted u_{rms}
- I don't like the formula:

$$\sqrt{\frac{3RT}{M}}$$

If use $R = 8.3145 \text{ J/mol}\cdot\text{K}$, get units of m/s



5-23

Gas Terms

Diffusion: describes the mixing of gases. The rate of diffusion is the rate of gas mixing.

Effusion: describes the passage of gas into an evacuated chamber.

5-24

Equations

$$\text{Effusion: } \frac{\text{Rate of effusion for gas 1}}{\text{Rate of effusion for gas 2}} = \frac{\sqrt{M_2}}{\sqrt{M_1}}$$

$$\text{Diffusion: } \frac{\text{Distance traveled by gas 1}}{\text{Distance traveled by gas 2}} = \frac{\sqrt{M_2}}{\sqrt{M_1}}$$

5-25

Sample Test Questions

- Ammonia and hydrochloric acid are both gases, which react with each other to form 'smoke' of ammonium chloride. If a sample of each is allowed to enter a glass tube 1.000 m long at the same time from opposite ends of the tube, where will the 'smoke' form?

5-26

Real Gases

Must correct ideal gas behavior when at high pressure (smaller volume) and low temperature (attractive forces become important).

5-27

Real Gases

$$[P_{\text{obs}} + a \left(\frac{n}{V}\right)^2] \times (V - nb) = nRT$$

$\underbrace{\quad\quad\quad}_{\text{corrected pressure}} \quad \quad \quad \underbrace{\quad\quad\quad}_{\text{corrected volume}}$
 $\uparrow \quad \quad \quad \quad \quad \quad \quad \uparrow$
 $P_{\text{ideal}} \quad \quad \quad \quad \quad \quad \quad V_{\text{ideal}}$

"Fudge Factors"

Table 5.4 Chang, 9th Ed.

TABLE 5.3
Values of the van der Waals
Constants for Some Common
Gases

Gas	$a \left(\frac{\text{atm} \cdot \text{L}^2}{\text{mol}^2} \right)$	$b \left(\frac{\text{L}}{\text{mol}} \right)$
He	0.0341	0.0237
Ne	0.211	0.0171
Ar	1.35	0.0322
Kr	2.32	0.0398
Xe	4.19	0.0511
H ₂	0.244	0.0266
N ₂	1.39	0.0391
O ₂	1.36	0.0318
Cl ₂	6.49	0.0562
CO ₂	3.59	0.0427
CH ₄	2.25	0.0428
NH ₃	4.17	0.0371
H ₂ O	5.46	0.0305

5-28

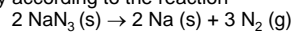
Sample Test Questions

- A balloon is filled to a volume of 5.00×10^2 mL at a temperature of 30.0 °C. The balloon is then cooled at constant pressure to a temperature of 1.00×10^2 K. What is the final volume of the balloon?
- If 25.0 mL of NO₂ gas is converted to N₂O₄ under constant conditions, what volume will the N₂O₄ occupy?
 $2 \text{NO}_2(\text{g}) \rightarrow \text{N}_2\text{O}_4(\text{g})$
- A student adds 2.50 g dry ice (solid CO₂) to an empty balloon. What will the volume be after the dry ice *sublimes*?

5-29

Sample Test Questions

- Air bags are activated when a severe impact causes a steel ball to compress a spring and electrically ignite a detonator cap. This causes sodium azide to decompose explosively according to the reaction



What mass of $\text{NaN}_3 (\text{s})$ must be reacted to inflate an air bag to 70.0 L at STP?

- A mixture of 1.00 g H_2 and 1.00 g He is placed in a 1.00-L container at 27 °C. Calculate the partial pressure of each gas and the total pressure.
- Calculate the average kinetic energy of the N_2 molecules in a sample of N_2 gas at 273 K and 546 K.

5-30

Sample Test Questions

- It took 4.5 minutes for 1.0 L helium to effuse through a porous barrier. How long will it take for 1.0 L Cl_2 gas to effuse under identical conditions?