CHAPTER 5: Gases

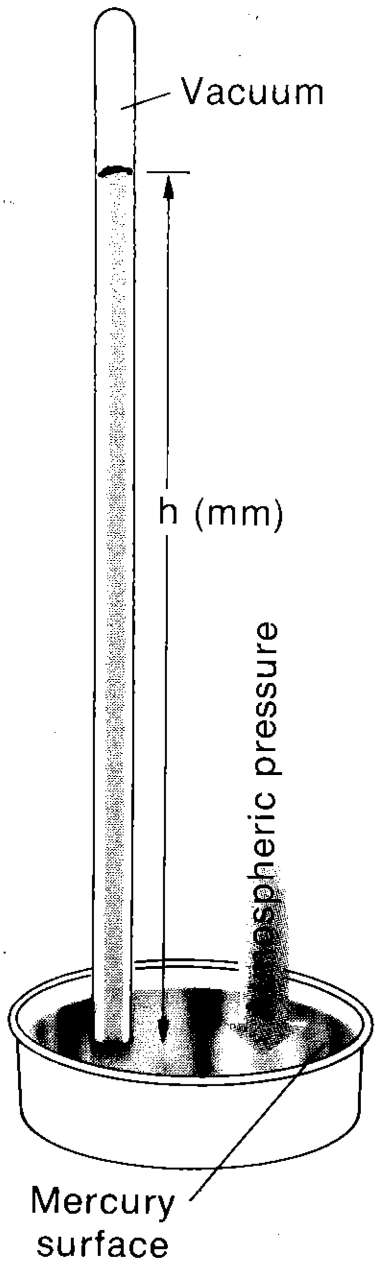
Section 5.1: Measurement on Gases Pressure, Volume and Temperature

Relationship Among Pressure Units

$$1 \text{ atm} = 760 \text{ mm Hg} = 14.70 \text{ lb/in}^2 = 1033 \text{ g/cm}^2 = 101.3 \text{ kPa}$$

$$1 \text{ torr} = 1 \text{ mm Hg}$$

$$D_{Hg} = 13.59 \text{ g/cm}^3$$



Relations Between Length, Volume, and Mass Units

METRIC		ENGLISH		METRIC-ENGLIGH	
Length					
1 km	= 10^3 m	1 ft	= 12 in	1 in	= 2.54 cm*
1 cm	= 10^{-2} m	1 yd	= 3 ft	1 m	= 39.37 in
1 mm	= 10^{-3} m	1 mile	= 5280 ft	1 mile	= 1.609 km
1 nm	= 10^{-9} m = 10Å				
Volume					
1 m^3	= 10^6 cm^3 = 10^3 L	1 gallon	= 4 qt = 8 pt	1 ft^3	= 28.32 L
1 cm^3	= 1 mL = 10^{-3} L	1 qt (Can.)	= 69.35 in^3	1 L	= 0.8799 qt (Can.)
		1 qt (U.S. liq.)	= 57.75 in^3	1 L	= 1.057 qt (U.S. liq.)
Mass					
1 kg	= 10^3 g	1 lb	= 16 oz	1 lb	= 453.6 g
1 mg	= 10^{-3} g	1 short ton	= 2000 lb	1 g	= 0.03527 oz
1 metric ton	= 10^3 kg			1 metric ton	= 1.102 short ton

*This conversion factor is exact; the inch is defined to be exactly 2.54 cm. The other factors listed in this column are approximate, quoted to four significant figures. Additional digits are available if needed for very accurate calculations. For example, the pound is defined to be 453.59237 g.

1. A five-gallon propane tank contains 0.784 mol of propane (C_3H_8) at 68°F. Express the volume of the tank in liters, the amount of propane in the tank in grams, and the temperature of the tank in Kelvin.

3. Complete the following table of pressure conversions.

mm Hg	Atmospheres	Kilopascals
913	———	———
———	0.833	———
———	———	122

CHAPTER 5: GASES

Section 5.2: The Gas Laws of Boyle, Charles, Gay-Lussac & Avogadro

Name the formulas of each gas law:

BOYLE:

CHARLES':

GAY-LUSSAC:

AVOGADRO'S:

1. A diver at a depth of 100 ft (pressure approximately 3 atm) exhales a small bubble of air with a volume equal to 100 mL. What will be the volume of the bubble (assume the same amount of air) at the surface?

2. What would be the volume of gas contained in an expandable 1.0 L cylinder at 15 Mpa (1 Mpa = 10^6 Pa) be at 1 atm (assuming constant temperature)?

3. A sample tube containing 103.6 mL of CO gas at 20.6 torr is connected to an evacuated 1.13-liter flask. (The new volume is the sum of those of the tube and the flask.) What will the pressure be when the CO is allowed into the flask?

4. A gas has a pressure of 3.2 atm and occupies a volume of 45 L. What will the pressure be if the volume is compressed to 27 L at a constant temperature?

5. The volume of a gas (held at constant pressure) is to be used "as a thermometer." If the volume at 0.0 °C is 75.0 cm^3 what is the temperature when the measured volume is 56.7 cm^3?

6. The gas in a closed container has a pressure of 3.00 x 10^2 Pa at 30 °C. What will the pressure be if the temperature is lowered to –172 °C?

7. The gas in a closed balloon has a pressure of 795 torr at 25 °C. What will the pressure be if the temperature is doubled? Is halved?

8. If a 16.6 L sample of a gas contains 9.2 moles of F_2, how many moles of gas would there be in a 750 mL sample at the same temperature and pressure?

9. An 11.2 L sample of gas is determined to contain 0.50 moles of N_2. At the same temperature and pressure how many moles of gas would there be in a 20. L sample?

10. Consider a 3.57 L sample of an unknown gas at a pressure of 4.3 x 10^3 Pa. If the pressure is changed to 2.1 x 10^4 Pa at a constant temperature, what will the new volume of the gas be?

Section 5.2/5.3: The Ideal Gas Law

Initial and Final States

12. A basketball is inflated in a garage at 25°C to a gauge pressure of 8.0 psi. Gauge pressure is the pressure above atmospheric pressure, which is 14.7 psi. The ball is used on the driveway at a temperature of −7°C and feels "flat." What is the actual pressure of the air in the ball? What is the gauge pressure?

14. A 3.50-cm^3 air bubble forms in a deep lake at a depth where the temperature is 6°C at a total pressure of 2.50 atm. The bubble rises to a depth where the temperature and pressure are 13°C and 1.75 atm, respectively. Assuming that the amount of air in the bubble has not changed, calculate its new volume.

Ideal Gas Law; Calculation of One Variable

20. Compressed-air tanks used by scuba divers have a volume of 8.0 L and are filled with air to a pressure of 135 atm at 20°C. How many grams of helium are required to fill a tank under the above conditions?

22. Complete the following table for carbon monoxide gas.

	Pressure	Volume	Temperature	Moles	Grams
a)	493 mm Hg	3.75 L	36°C	———	———
b)	1.28 atm	6.39 L	———	0.500	———
c)	125 kPa	———	99°C	———	43.2
d)	———	2.98 L	125°C	———	0.827

Ideal Gas Law; Density and Molar Mass

25. Calculate the densities (in g/L) of the following gases at 97°C and 755 mm Hg.

(a) hydrogen chloride

(b) sulfur dioxide

29. Freon is a gas made up of carbon, fluorine, and chlorine atoms. It was used as a refrigerant in car air conditioners. It is also one of the culprits in the depletion of the ozone layer. It has a density of 4.65 g/L at 735 mm Hg and 33°C.

(a) What is the molar mass of Freon?

(b) Freon is made up of 9.92% C, 58.6% Cl, and 31.4% F. What is its molecular formula?

30. Phosgene is a highly toxic gas made up of carbon, oxygen, and chlorine atoms. Its density at 1.05 atm and 25°C is 4.24 g/L.

(a) What is the molar mass of phosgene?

(b) Phosgene is made up of 12.1% C, 16.2% O, and

33. A 1.58-g sample of $C_2H_3X_3(g)$ has a volume of 297 mL at 769 mm Hg and 35°C. Identify the element X.

CHAPTER 5: Gases

Section 5.4: Stoichiometry of Gaseous Reactions

The first stoichiometric relationship to de discovered was the **law of combining volumes** proposed by Gay-Lussac in 1808:

The volume ratio of any two gases in a reaction at constant temperature is the same as the reacting mole ratio.

These integer volume ratios confirm the proposal first made in 1811 by an Italian scientist, Amadeo Avogadro. It is called **Avogadro's hypothesis.** It has been used successfully in explaining the properties of gases for almost two centuries.

Equal volumes of gases, measured at the same temperature and pressure, contain equal numbers of molecules.

Gases in Reactions

35. Dichlorine oxide is used as a bacteriocide to purify water. It is produced by the reaction

$$SO_2(g) + 2Cl_2(g) \longrightarrow SOCl_2(l) + Cl_2O(g)$$

How many liters of dichlorine oxide can be produced from 6.0 L of chlorine? Assume 100% yield and that all the gases are measured at the same temperature and pressure.

36. Hydrogen sulfide gas (H_2S) is responsible for the foul odor of rotten eggs. When it reacts with oxygen, sulfur dioxide gas and steam are produced.

(a) Write a balanced equation for the reaction.

(b) How many liters of oxygen would be required to react with 12.0 L of hydrogen sulfide? Assume 100% yield and constant temperature and pressure.

37. Hydrogen cyanide, HCN, is a poisonous gas that is used in the gas chamber for the execution of people sentenced to death. It can be formed by the reaction

$$NaCN(s) + H^{+}(aq) \longrightarrow HCN(g) + Na^{+}(aq)$$

What volume of 6.00 M HCl is required to react with excess NaCN to produce enough HCN to fill a room with volume 27 m^3 at a pressure of 0.987 atm and 23°C?

a) Using the Ideal Gas Law, compute the ~~volume~~ moles of $HCN_{(g)}$ to be produced.

b) From the balanced equation, compute the moles of $H^{+1}{}_{(aq)}$ required.

c) Knowing the moles of $H^{+1}{}_{(aq)}$ and the molarity of the HCl solution, compute the volume of HCl to be used.

Dinitrogen oxide, commonly called nitrous oxide, is used as a propellant gas for whipped-cream dispensers. It is prepared by heating ammonium nitrate to 250°C. Water vapor is also formed.

(a) Write a balanced equation for the decomposition of ammonium nitrate into nitrous oxide and steam.

(b) What volume of dinitrogen oxide gas is formed at 250°C and 1.0 atm when 5.00 g of NH_4NO_3 is heated? Assume 100% yield.

i. Compute the moles of $NH_4NO_{3(s)}$ to be reacted.

ii. From the balanced equation, compute the moles of dinitrogen $oxide_{(g)}$ formed.

iii. Using the Ideal Gas Law, compute the volume of dinitrogen oxide gas formed.

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Section 5.5 Gas Mixtures: Partial Pressures and Mole Fractions

In 1801, John Dalton proposed a relationship that applies to gas mixtures. For a mixture of two gases A and B, each gas would exert a partial pressure equal to the pressure it would exert as if it occupied the entire volume by itself. This is referred to as **Dalton's Law of Partial Pressures:**

The total pressure of a gas mixture is the sum of the partial pressures of the components of the mixture.

$$P_{tot} = P_A + P_B$$

A relationship commonly used to compute partial pressures of gases in a mixture when the total pressure and the composition of the mixture are known is as follows:

The partial pressure of a gas in a mixture is equal to its mole fraction multiplied by ithe total pressure.

1. Mole fraction
   For gas A, it is the fraction of the total number of moles that is accounted for by gas A.

$$X_A = \text{mole fraction of gas A} = \frac{n_A}{n_{total}}$$

n_A = number of moles of gas A

n_{total} = number of moles of all gases combined

2. Mathematical relationship between the mole fraction and partial pressure

$$P_A = X_A P_{total}$$

41. Some chambers used to grow bacteria that thrive on CO_2 have a gas mixture made up of 95.0% CO_2 and 5.0% O_2 (mole percent). What is the partial pressure of each gas if the total pressure is 735 mm Hg?

A sample of a smoke stack emission was collected into a 1.25-L tank at 752 mm Hg and analyzed. The analysis showed 92% CO_2, 3.6% NO, 1.2% SO_2, and 4.1% H_2O by mass. What is the partial pressure exerted by each gas?

Wet Gases: Partial Pressure of Water.

When a gas (such as hydrogen) is collected by bubbling through water, it picks up water vapor; molrcules of water escape from the liquid and are carried along with the gas. Dalton's law can be applied to the resulting gas mixture:

$$P_{tot} = P_{H_2O} + P_{gas}$$

In this case, P_{tot} is the measured pressure. The partial pressure of water vapor, P_{H_2O}, is equal to the ***vapor pressure*** of liquid water. It has a fixed value at a given temperature (see below).

Vapor Pressure of Water at Various Temperatures

Temperature (°C)	Pressure (mmHg)	Temperature (°C)	Pressure (mmHg)
15	12.8	23	21.0
16	13.6	24	22.4
17	14.5	25	23.7
18	15.5	26	25.2
19	16.5	27	26.7
20	17.5	28	28.3
21	18.6	29	30.0
22	19.8	30	31.8

) A student collects 355 cm^3 of oxygen saturated with water vapor at 27°C. The mixture exerts a total pressure of 775 mm Hg. At 27°C, the vapor pressure of $H_2O(l)$ = 26.7 mm Hg.

a. What is the partial pressure of oxygen in the sample?

b. How many grams of oxygen does the sample contain?

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Section 5.6: Kinetic Theory of Gases

The fact that the ideal gas applies to all gases indicates that all gases must have common properties. Between about 1850 and 1880, James Maxwell, Rudolf Clausius, Ludwig Boltzmann (and others) developed the **Kinetic Theory of Gases**.

Postulates of the Kinetic Theory

The kinetic theory of gases is based on several assumptions, including the following.

1. Gases consist of atoms or molecules in continuous, random motion. These particles undergo frequent collisions with each other and with the walls of their container. Gas pressure is caused by collisions with the walls.
2. Collisions between gas particles are elastic; there is no change in total energy when a collision occurs. No kinetic energy is converted to heat, which explains why the temperature of an insulated gas does not change with time.
3. The volume occupied by gas particles is negligibly small compared with that of their container. This assumption, like (4) below, is valid at ordinary temperatures and pressures, where a gas is mostly "empty space."
4. Attractive forces between particles have a negligible effect on their behavior. The atoms or molecules in a gas can be treated as independent particles.

The two most important postulates for our purposes are:

5. ***The average translational kinetic energy, E_t, of a gas particle is directly proportional to the absolute temperature.*** That is,

$$mu^2 = cT \; ; \; E_t = mu^2/2$$

where c is a constant.

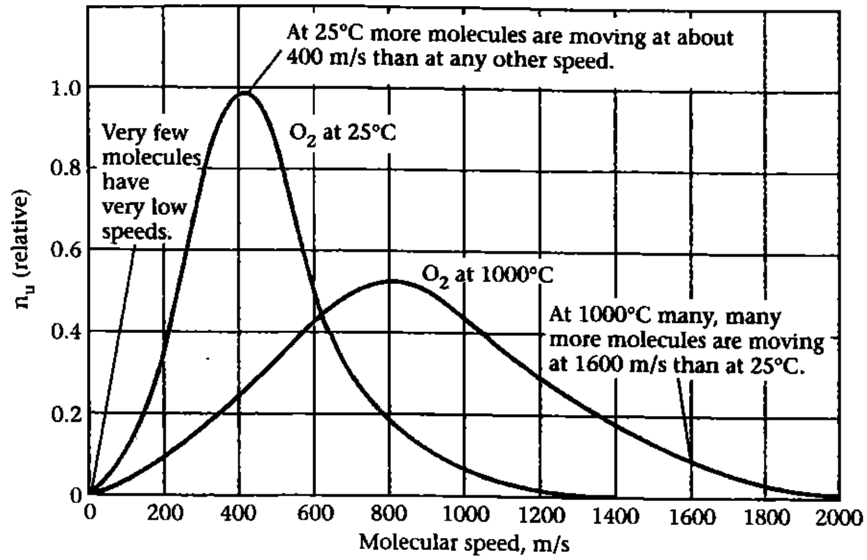
The constant c in this equation can be evaluated from kinetic theory; it turns out to be

$$c = 3R$$

*More rigorously, u^2 is the average of the squares of the speeds of all molecules.

Distribution of Molecular Speeds and Energies

Figure 5.6
The distribution of molecular velocities in oxygen gas at two different temperatures, 25°C and 1000°C. At the higher temperature, the fraction of molecules moving at very high speeds is much greater.



As temperature increases, the speed of the molecules increases. The distribution curve for molecular speeds (Fig. 5.6) shifts to the right and becomes broader. The chance of a molecule having a very high speed is much greater at 1000°C than it is at 25°C. Note, for example, that a large number of molecules have speeds greater than 1200 m/s at 1000°C.

Average Speeds of Gas Particles

1) Compute the average speed of a nitrogen molecule at 25°C.

2) A professional tennis player can serve a tennis ball at 45.0 m/s. At what temperature will a helium atom have the same average speed?

Effusion of Gases; Graham's Law

At a given temperature and pressure, the rate of effusion of a gas is inversely proportional to the square root of its molar mass.

One way to check the validity of calculations made from kinetic theory is to study the process of **effusion,** the flow of gas particles through tiny pores or pinholes. The relative rates of effusion of different gases depend upon two factors: the pressures of the gases and the relative speeds of their particles. If two different gases A and B are compared at the same pressure, only their speeds are of concern, and

$$\frac{\text{rate of effusion B}}{\text{rate of effusion A}} = \frac{u_B}{u_A}$$

where u_A and u_B are average speeds.

45. A gas effuses 1.55 times faster than propane (C_3H_8) at the same temperature and pressure.

(a) Is the gas heavier or lighter than propane?

(b) What is the molar mass of the gas?

46. What is the ratio of the rate of effusion of the most abundant gas, nitrogen, to the lightest gas, hydrogen?

50. It takes 12.6 s for 1.73×10^{-3} mol of CO to effuse through a pinhole. Under the same conditions, how long will it take for the same amount of CO_2 to effuse through the same pinhole?

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Section 5.7 Real Gases: The Van der Waals Equation

A careful study of deviations from ideal behavior led the Dutch scientist J. D. van der Walals to propose a modified version of the (P)(V) relation. His modification is based on the observation that observed pressures generally are smaler than predicted by Boyle's law and observed volumes are generally greater than so predicted,

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

The correction term n^2a/V^2 is added to the observed pressure and the correction term nb is subtracted from the observed volume. The number of moles of gas is given by n; a and b are constants characteristic of the gas in question.

Real Gases

Calculate the pressure of one mole of oxygen gas in 125-mL container at 0°C using the

a. Ideal Gas Law.

b. van der Waals equation (see Table 4.3).

TABLE 4.3 van der Waals Constants

GAS	a $\left(\frac{L^2 \cdot atm}{mol^2}\right)$	b (L/mol)
H_2	0.244	0.027
O_2	1.360	0.032
N_2	1.390	0.039
CH_4	2.253	0.043
CO_2	3.592	0.043
SO_2	6.714	0.056
Cl_2	6.493	0.056
H_2O	5.464	0.030