Chapter 11
Outline

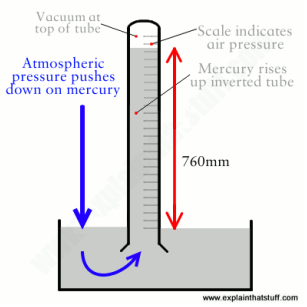
Main Ideas

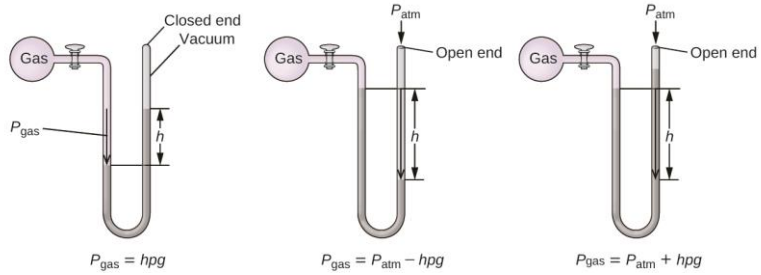
I. Gases & Pressure

A. Pressure & Force

1. Collisions of gas molecules cause an outward push, or force, any surface contacted.
   a. *Pressure* = force/area = Newtons/ cm^2
   b. *Newton* = (kilogram)(meter)/ $second^2$
   c. Earth's gravity has acceleration of 9.8 m/s^2, so weight equals mass x 9.8 m/s^2
   d. Air exerts pressure; at sea level, atmospheric pressure = 1.03 kg/cm^2 = 10.1 N/cm^2 = 1.013 kPa

2. Measuring Pressure
   a. *Barometer* measures atmospheric pressure by balancing the weight of a liquid with weight of air.
      i. Hg at 1 atmosphere of pressure can only rise to 760 mm, regardless of the size of the tube.
   b. Manometer measures pressure of gas sample by comparing heights of Hg in 2 arms of a U-tube.





3. Units of Pressure
   a. Millimeters of mercury (mm Hg)
   b. Torr; 1 torr ≈ 1 mm Hg
   c. Atmosphere of pressure (atm)
   d. Pascal (Pa) = 1 N/ m^2
   e. Pounds per square inch (psi)
   f. Equivalents: 1 atm = 760 mm Hg = 760 torr = 101.325 kPa = 14.70 psi

4. Standard Temperature and Pressure [STP]
   a. STP = 1 atm at 0°C; 0 °C = 273.15 Kelvin (K = 273 + °C)

B. Dalton's Law of Partial Pressure

1. Partial pressure = pressure of each gas in a mixture

2. Total pressure = the sum of all partial pressures of the component gases; $P_T = P_1 + P_2 + P_3 + \ldots$
   a. Per the kinetic-molecular theory, each gas exerts a pressure independent of that from other gases

3. Gases Collected by H_2O Displacement
   a. Gases produced in the lab are often collected over water and will displace that water
   b. Gases collected over water are mixed with water molecules which exert *vapor pressure*
   c. $P_{atm} = P_{gas} + P_{water}$

II. The Gas Laws

A. Boyle's Law: Pressure-Volume Relationship

1. At a constant temperature & # moles of gas, as volume increases, pressure decreases.

2. $PV = k$: $P_1V_1 = P_2V_2$

3. Per kinetic-molecular theory: pressure increases with increased # of collisions. As volume increases, there are less molecules per unit volume and therefore, less collisions, less pressure.

B. Charles's Law: Volume-Temperature Relationship
1. At a constant pressure, volume increases as Kelvin temperature increases.

2. $V = kT$: $V_1/T_1 = V_2/T_2$: $V_1T_2 = V_2T_1$

C. Gay-Lussac's Law: Pressure-Temperature Relationship
1. At a constant volume of a fixed quantity of gas, pressure is directly proportional to temperature.

2. $P = k\,T$: $P_1/T_1 = P_2/T_2$: $P_1T_2 = P_2T_1$

D. Combined Gas Laws
1. $PV/T = k$: $[(P_1V_1) / T_1] = [P_2V_2) / T_2]$

III. Gas Volumes & the Ideal Gas Law
A. Measuring & Comparing Volumes of Reacting Gases
1. *Gay-Lussac's Law of combining volumes of gases*
a. Volumes of reactants & products can be expressed as ratios of small whole numbers
b. Example: 2 volumes of $H_{2\,(g)}$ + 1 volume of $O_{2\,(g)} \rightarrow$ 2 volumes of $H_2O_{\,(g)}$

B. Avogadro's Law
1. Rejected Dalton's theory about gaseous elements existing only as isolated single atoms

2. Proposed that equal volumes of gases at same temp. & pressure have equal # of molecules

3. $V = kn$, where V = volume, n = # moles, & k is a constant

4. $H_{2\,(g)} + Cl_{2\,(g)} \rightarrow 2\,HCl_{\,(g)}$ can be seen as 1 volume/molecule + 1 volume/molecule $\rightarrow$ 2 volumes/molecules

C. Molar Volume of a Gas
1. 1 mole of a substance contains 6.022×10^{23} particles or atoms or molecules
a. 1 mole of O_2 has 6.022×10^{23} diatomic oxygen molecules with a mass of 31.9988 g
b. 1 mole of He has 6.022×10^{23} monatomic helium molecules with a mass of 4.002602g

2. *Standard molar volume of a gas* is volume of 1 mole of gas at STP & equals **22.41410 L**

D. Gas Stoichiometry
1. For gaseous reactions, coefficients indicate molar amounts & mole ratios & volume ratios

E. The Ideal Gas Law
1. $PV = nRT$ where P = pressure, V = volume, n = # moles, R = gas constant, & T = temperature

F. The Ideal Gas Constant
1. Units depend on units of pressure used, i.e., atm, kPa, mm Hg, psi, etc…
a. 0.082058 L•atm/ mol•K
b. 8.314 L•kPa/ mol•K
c. 62.4 L•mm Hg/ mol•K

IV. Diffusion and Effusion
A. Graham's Law of Effusion
1. $KE = \frac{1}{2}\,mv^2$

2. $\frac{1}{2}\,M_A\,v_A^{\,2} = \frac{1}{2}\,M_B\,v_B^{\,2}$

3. (rate of effusion of A) / (rate of effusion of B) = $\sqrt{M_B} / \sqrt{M_A}$