Behavior of Gases

Properties of Gases
- Gases have weight
- Gases take up space
- Gases exert pressure
- Gases fill their containers
- Gases are mostly empty space
- The molecules in a gas are separate, very small and very far apart

Kinetic Theory of Matter:
- Gas molecules are in constant, chaotic motion
- Collisions between gas molecules are elastic (there is no energy gain or loss)
- The average kinetic energy of gas molecules is directly proportional to the absolute temperature
- Gas pressure is caused by collisions of molecules with the walls of the container

Measurement of Gases

To describe a gas, its volume, amount, temperature, and pressure are measured.

- **Volume**: measured in L, mL, cm^3 ($1\ mL = 1\ cm^3$)
- **Amount**: measured in moles (mol), grams (g)
- **Temperature**: measured in Kelvin (K) $K = °C + 273$
- **Pressure**: measured in mm Hg, torr, atm, etc.
  $P = F/A$ (force per unit area)

Units of Pressure:

1 atmosphere = 760 mm Hg
1 atm = 760 torr
1 atm = 1.013×10^5 Pa
1 atm = 101.3 kPa
1 atm = 1.013 bar

$PV = nRT$

$$\frac{P_1V_1}{n_1RT_1} = \frac{P_2V_2}{n_2RT_2}$$

Intro to Gas Laws:

Boyle's Law: relation of volume to pressure

$V \propto 1/P$

$P_1V_1 = P_2V_2$

Example: A sample of gas occupies 12 L under a pressure of 1.2 atm. What would its volume be if the pressure were increased to 3.6 atm? (assume temp is constant)

$(1.2\ atm)(12\ L) = (3.6\ atm)V_2$

$V_2 = 4.0\ L$

Charles' Law: relation of volume to temperature

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

**temp must be expressed in Kelvin!

Example: A sample of nitrogen gas occupies 117 mL at 100°C. At what temperature would it occupy 234 mL if the pressure does not change? (express answer in K and °C)

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

$$\frac{117\ ml}{373\ K} = \frac{234}{T_2} = \boxed{746\ K}$$

$746 - 273 = \boxed{473\ °C}$

Standard Temperature & Pressure (STP): 0°C (273 K) and 1 atm (760 torr, 760 mm Hg)

The Combined Gas Law Equation

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2} \quad \text{or} \quad P_1V_1T_2 = P_2V_2T_1$$

Examples:

1. A sample of neon gas occupies 105 L at 27°C under a pressure of 985 torr. What volume would it occupy at standard conditions?

$$\frac{(985\ torr)105\ L}{300\ K} = \frac{(760\ torr)V_2}{273\ K}$$

2. A sample of gas occupies 10.0 L at 240°C under a pressure 80.0 kPa. At what temperature would the gas occupy 20.0 L if we increased the pressure to 107 kPa?

$P_1V_1T_2 = P_2V_2T_1$

$(80.0\text{ kPa})(10.0\text{ L})(T_2) = (107\text{ kPa})(20.0\text{ L})(513\text{ K})$

$T_2 = 1372\text{ K} \approx \boxed{1370\text{ K}}$

3. A sample of oxygen gas occupies 23.5 L at 22.2°C and 1.3 atm. At what pressure (in mm Hg) would the gas occupy 11.6 L if the temperature were lowered to 12.5°C?

$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$ $\quad \frac{(988\text{ mm})(23.5\text{ L})}{22.2°C+273} = \frac{P_2(11.6\text{ L})}{12.5°C+273}$

$\frac{(988\text{ mm})(23.5\text{ L})(285\text{ K})}{3424.32} = \frac{P_2(11.6\text{ L})(295.2\text{ K})}{3424.32}$

$P_2 = 1932.4 \approx \boxed{1900\text{ mmHg}}$

Gases: Standard Molar Volume & The Ideal Gas Law

· **Avogadro's Law**: at the same temperature and pressure, equal volumes of all gases contain the same # of molecules (moles).

· **Standard Molar Volume** = 22.4 L @ STP

(this is true of "ideal" gases; at reasonable temperatures & pressures, the behavior of many "real" gases is nearly ideal)

Example: 1.00 mole of a gas occupies 36.5 L, and its density is 1.36 g/L at a given temperature & pressure.

a) What is its molecular weight? (molar mass)

$\frac{1.36\text{ g}}{\text{L}} \times \frac{36.5\text{ L}}{1.00\text{ mol}} = 49.6\text{ g/mol}$

b) What is the density of the gas at standard conditions?

$\frac{49.6\text{ g}}{\cancel{\text{mol}}} \times \frac{1\cancel{\text{mol}}}{22.4\text{ L}} = \boxed{2.21\text{ g/L}}$

The Ideal Gas Law: shows the relationship among the pressure, volume, temp., and the # of moles in a sample of gas.

Where,
- P = pressure (atm)
- V = volume (L)
- n = # moles
- T = temp (K)
- R = universal gas constant
- = $0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}$

(the units of R depend on the units used for P, V, and T)

Examples:

1) What volume would 50.0 g of ethane, C_2H_6, occupy at 140°C under a pressure of 1820 torr?

$1820 \times \frac{1\ atm}{760\ torr} = 2.39\ atm$

$50.0\ g\ C_2H_6 \times \frac{1\ mol\ C_2H_6}{30.08\ g\ C_2H_6} = 1.662\ mol$

$PV = nRT$

$(2.39\ atm)(V) = (1.662\ mol)(0.0821)(140°C + 273)$

$\boxed{V = 23.6\ L}$

2) Calculate: a) the # of moles in, and (b) the mass of an 8.96 L sample of methane, CH_4, measured at standard conditions.

$8.96\ L\ CH_4 \times \frac{1\ mol\ CH_4}{22.4\ L\ CH_4} = \boxed{0.400\ mol\ CH_4}$

$0.400\ mol\ CH_4 \times \frac{16.05\ g\ CH_4}{1\ mol} = \boxed{6.42\ g\ CH_4}$

3) Calculate the pressure exerted by 50.0 g of ethane, C_2H_6, in a 25.0 L container at 25°C.

1.662 mol

$\frac{P(25.0\ L)}{25.0\ L} = \frac{(1.662\ mol)(0.0821)(25°C + 273)}{25.0\ L}$

$\boxed{P = 1.63\ atm}$

Determining Molecular Weights & Molecular Formulas of Gases: if the mass of a volume of gas is known, we can use this info. to determine the molecular formula for a compound.

Examples:

1) A 0.109 g sample of a pure gaseous compound occupies 112 mL at 100°C and 750 torr. What is the molecular weight of the compound?

$PV = nRT$

$(0.99\ atm)(0.112\ L) = (n)(0.0821)(100°C + 273K)$

$n = 0.00362\ mol$

$1\ mol = \frac{0.109\ g}{0.00362\ mol} = \boxed{30.1\ g}$

2) A compound that contains only C and H is 80.0% C and 20.0% H by mass. At STP 546 mL of the gas has a mass of 0.732 g. What is the molecular (true) formula for the compound?

$80.0\ g\ C \times \frac{1\ mol}{12.0\ g} = \frac{6.67\ mol\ C}{6.67\ mol} = 1$

$20.0\ g\ H \times \frac{1\ mol}{1.0\ g} = \frac{20.0\ mol\ H}{6.67\ mol} = 3$

Empirical Formula $= CH_3$

$n = 0.546\ L \times \frac{1\ mol}{22.4\ L} = 0.024\ mol$

$Molar\ Mass = \frac{0.732\ g}{0.024\ mol} = 30.0\ g/mol$

$\frac{30.0\ g/mol}{15.04\ g/mol} = 2$

$CH_3 \times 2 = \boxed{C_2H_6}$

Notes: Partial Pressures and mole Fraction

In a mixture of gases each gas exerts the pressure it would exert if it occupied the volume alone.

The total pressure exerted by a mixture of gasses is the sum of the partial pressures of the individual gases:

$$P_{total} = P_1 + P_2 + P_3 + \ldots$$

Example: If 100.0 mL of hydrogen gas, measured at 25°C and 3.00 atm, and 100.0 mL of oxygen, measured at 25°C and 2.00 atm, what would be the pressure of the mixture of gases?

$$P_T = 3.00 \text{ atm} + 2.00 \text{ atm}$$
$$P_T = 5.00 \text{ atm}$$

Vapor Pressure of a Liquid:

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

or

$$P_{gas} = P_{atm} - P_{H_2O}$$

Temp (°C)	v.p. of water (mm Hg)	Temp. (°C)	v.p. of water (mm Hg)
18	15.48	21	18.65
19	16.48	22	19.83
20	17.54	23	21.07

Example: A sample of hydrogen gas was collected by displacement of water at 25°C. The atmospheric pressure was 748 mm Hg. What pressure would the dry hydrogen exert in the same container?

$$P_{H_2} = P_{atm} - P_{H_2O}$$
$$P_{H_2} = 748 \text{ mm Hg} - 23.76 \text{ mm Hg}$$
$$P_{H_2} = 724.24 \text{ mm Hg}$$

Example: A sample of oxygen was collected by displacement of water. The oxygen occupied 742 mL at 27°C. The barometric pressure was 753 mm Hg. What volume would the dry oxygen occupy at STP?

$P_{O_2} = P_{atm} - P_{H_2O}$

$P_{O_2} = 753\ mm\ Hg - 26.74\ mm\ Hg$

$P_{O_2} = 726\ mm\ Hg$

$P_1V_1T_2 = P_2V_2T_1$

$(726\ mm\ Hg)(0.742\ L)(273\ K) = (760\ mm\ Hg)(V_2)(300\ K)$ $V_2 = 0.645\ L$

Example: A student prepares a sample of hydrogen gas by electrolyzing water at 25°C. She collects 152 mL of H_2 at a total pressure of 758 mm Hg. Calculate (a) the partial pressure of hydrogen, and (b) the number of moles of hydrogen collected.

a.) $P_{H_2} = P_{atm} - P_{H_2O}$

$P_{H_2} = 758\ mm\ Hg - 23.76\ mm\ Hg$

$= 734\ mm\ Hg$

b) $PV = nRT$

$(0.966\ atm)(0.152\ L) = (n)(0.0821)(298\ K)$

$\boxed{n = 0.00600\ mol\ H_2}$

Graham's Law of Diffusion & Effusion

$$\frac{rate_1}{rate_2} = \sqrt{\frac{MM_2}{MM_1}}$$

Where,

Rate = rate of diffusion or effusion

MM=molar mass

Stoichiometry of Gaseous Reactions

A balanced equation can be used to relate moles or grams of substances taking part in a reaction. (AND VOLUME!)

Example: Hydrogen peroxide is the active ingredient in commercial preparations for bleaching hair. What mass of hydrogen peroxide must be used to produce 1.00 L of oxygen gas at 25 °C and 1.00 atm?

$2\ mol\ H_2O_2 = 1\ mol\ O_2$

$2H_2O_2 \rightarrow O_2 + 2H_2O$

$T_{O_2} = 25°C = 298\ K$

$P_{O_2} = 1.00\ atm$

$V_{O_2} = 1.00\ L$

$PV = nRT$

$(1.00\ atm)(1.00\ L) = (n)(0.0821)(298\ K)$

$n = 0.0409\ mol$

$0.0409\ mol\ mol\ O_2$

$0.0409\ mol\ O_2 \times \frac{2\ mol\ H_2O_2}{1\ mol\ O_2} \times \frac{34.02\ g\ H_2O_2}{1\ mol\ H_2O_2} = \boxed{2.78\ g\ H_2O_2}$