Gases

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1 Gases and Kinetic Theory

The following are a list of definitions in Gases and Kinetic Theory:

Gas - The state of matter that has the weakest attraction forces among its components (atoms or molecules). It occupies the whole volume of its container. Gases expand and/or contract with temperature and pressure

Pressure - The distribution of a force over a given area. Or Force per unit of area. SI unit is the Pascal, but in Chemistry we use atmospheres (atm.)

Barometer - Measures the pressure of air. (Developed by Torricelli in 1643). Other pressures by manometers. 1 standard atm = 760 mm Hg

Pascal - (Pa) SI Unit 1 atm = 101.325 katm

Properties of Gases - Can be modeled using math. Depends on V = Volume of gas(L) and T = Temperature (K) and n = amount (moles)
Directly Proportional - If one quantity is increased or decreased then the other quantity increases or decreases.
Indirectly Proportional - If one quantity increases then the other quantity decreases and vice-versa.
Avogadro's Hypothesis - V and n are directly proportional at constant P and T.
Diffusion - The movement of gases through each other (e.g. filling a balloon).
Effusion - The movement of gas through a tiny hole (e.g. poking a hole in a balloon).

2 Boyle's Law

1. P and V are inversely proportional when n and T are constant.
2. Must convert temperatures to Kelvin

$$Boyle'sLaw = Volume * Pressure = Constant \tag{1}$$

$$V_1P_1 = P_2V_2 \tag{2}$$

3 Charles' Law

1. V and T are directly proportional when n and P are constant.
2. Must convert temperatures to Kelvin

$$Charles'Law = \frac{Volume}{Temperature} = Constant \tag{3}$$

$$\frac{V_1}{T_1} = \frac{V_2}{T_2} \tag{4}$$

4 Combined Gas Law

1. Must convert temperatures to Kelvin

$$CombinedGasLaw = \frac{Volume * Pressure}{Temperature} = Constant \tag{5}$$

$$\frac{V_1P_1}{T_1} = \frac{V_2P_2}{T_2} \tag{6}$$

5 Ideal Gas Law

1. Pressure must always be in atmospheres (atm) for Ideal Gas Law
2. Must convert temperatures to Kelvin

$$IdealGasLaw = (Volume * Pressure) = n * R * T \tag{7}$$

$$VP = nRT \tag{8}$$

6 Universal Gas Constant

$$R = 0.08206\frac{L * atm}{mol * K} \tag{9}$$

7 Standard Temperature and Pressure (STP)

$$273K, 0°C, 1atm \tag{10}$$

8 Pressure Conversions

$$1atm = 760Torr = 760mmHg = 1.013bar = 101.3kPa = 14.7psi \tag{11}$$

9 Ideal Gas Law Applications

9.1 Density of a Gas

$$R = \text{Universal Gas Constant} \tag{12}$$

$$M = \text{Molar Mass} \tag{13}$$

$$Density = \frac{Pressure * MolarMass}{R * Temperature} \tag{14}$$

$$\rho = \frac{PM}{RT} \tag{15}$$

9.2 Molar Mass

$$\text{Molar Mass} = \frac{\text{mass in grams}}{\text{number of moles}} \tag{16}$$

$$M = \frac{m}{n} \tag{17}$$

$$n = \frac{m}{M} \tag{18}$$

Substitute the n we derived from molar mass into ideal gas law.

$$VP = nRT \tag{19}$$

$$VP = \frac{m}{M}RT \tag{20}$$

$$MVP = mRT \tag{21}$$

$$M = \frac{mRT}{VP} \tag{22}$$

10 Dalton's Law

10.1 Dalton's Law

$$P_{Total} = P_a + P_b + P_c + ... \tag{23}$$

10.2 Mole Fraction (X_A)

$$X_A = \frac{MolX}{TotalMol} \tag{24}$$

10.3 Partial Pressure (P_A)

$$P_A = (X_a)(P_{Total}) \tag{25}$$

11 Graham's Law

$$\frac{\text{Rate A}}{\text{Rate B}} = \sqrt{\frac{\text{Molar Mass B}}{\text{Molar Mass A}}} \tag{26}$$

$$\frac{R_1}{R_2} = \sqrt{\frac{M_2}{M_1}} \tag{27}$$

12 Gas Law Practice

1. A gas a pressure of 3800 mmHg what is the pressure in units of atm and bar?

$$\frac{3800mmHg}{760mmHg} = 5atm \tag{28}$$

$$5atm * 1.013bar = 5.065bar \tag{29}$$

2. A gas has a pressure of 2.5 atm and a volume of 33L. What will be the volume of the gas when the pressure increases to 7.5 atm (while the temperature remains constant)?

$$P_1V_1 = P_2V_2 \tag{30}$$

$$2.5atm * 33L = 7.5atm * V_2 \tag{31}$$

$$\frac{2.5atm * 33L}{7.5atm} = V_2 \tag{32}$$

$$11L = V_2 \tag{33}$$

3. A gas occupies 900.0 mL at a temperature of 27.0°C. What is the volume at 132.0°C?

$$\frac{V_1}{T_1} = \frac{V_2}{T_2} \tag{34}$$
$$T_1 = 27°C + 273K = 300K \tag{35}$$
$$T_2 = 132°C + 273K = 405K \tag{36}$$
$$\frac{900mL}{300K} = \frac{V_2}{405K} \tag{37}$$
$$1,215mL = V_2 \tag{38}$$
$$1,215mL * .001 = 1.215L = V_2 \tag{39}$$

4. A gas has an initial volume of 2.5 L and an initial temperature of -20°C. What will the final pressure be if the volume increases to 12.0 L and the temperature increases to 65°C?

$$\frac{V_1P_1}{T_1} = \frac{V_2P_2}{T_2} \tag{40}$$
$$\frac{2.5L * 0.134atm}{253K} = \frac{12.0L * P_2}{338K} \tag{41}$$
$$P_2 = 0.03729578atm \approx 0.0373atm \tag{42}$$

5. How many moles of gas occupy 98 L at a pressure of 2.8 atmospheres and a temperature of 292 K?

$$VP = nRT \tag{43}$$
$$98L * 2.8atm = n * (0.08206\frac{L * atm}{mol * K}) * 292K \tag{44}$$
$$\frac{98L * 2.8atm}{(0.08206\frac{L*atm}{mol*K}) * 292K} = n \tag{45}$$
$$n = 11.45169 \text{ mols} \approx 11.45 \text{ mols} \tag{46}$$

6. What is the density of N_2O that produces 5.0 L N_2O at 3.5 atm with T = 32°C?

$$N_2 = 14(2) = 28\frac{g}{mol} \tag{47}$$
$$O = 16(1) = 16\frac{g}{mol} \tag{48}$$
$$M = N_2 + O = 44\frac{g}{mol} \tag{49}$$

$$\rho = \frac{PM}{RT} \tag{50}$$

$$\rho = \frac{(3.5\text{atm})(44\frac{g}{mol})}{(0.08206\frac{\text{L*atm}}{\text{mol*K}})(305K)} = 6.15\frac{g}{L} \tag{51}$$

7. What is the molar mass of 95 grams of an unknown gas at 790 mmHg and $35°C$ at a Volume of 5L?

$$\frac{790mmHg}{760} = 1.039atm \tag{52}$$

$$M = \frac{mRT}{VP} \tag{53}$$

$$M = \frac{(95\text{g})(0.08206\frac{\text{L*atm}}{\text{mol*K}})(308K)}{(5L)(1.04atm)} = 462.2\frac{g}{mol} \tag{54}$$

8. A mixture of hydrogen gas and oxygen gas exerts a total pressure of 3.5 atm on the walls of its container. If the partial pressure of hydrogen is 2.0 atm, find the mole fraction of oxygen in the mixture.

$$P_A = X_A(P_{Total}) \tag{55}$$

$$2.0atm = X_A(3.5atm) \tag{56}$$

$$\frac{2.0atm}{3.5atm} = X_A = 0.5714mol \tag{57}$$

9. A 5 L flask contains 25 g of Cl_2, 73 g of O_2, and 50 g of N_2. What is the partial pressure of each gas if the total is pressure is 755 mmHg?

$$\frac{755mmHg}{760} = 0.993atm \tag{58}$$

$$Cl_2 = 35.453(2) = 70.906\frac{g}{mol} \tag{59}$$

$$O_2 = 16(2) = 32\frac{g}{mol} \tag{60}$$

$$N_2 = 14(2) = 28\frac{g}{mol} \tag{61}$$

$$Cl_2 + O_2 + N_2 = 130.906\frac{g}{mol} \tag{62}$$

$$\frac{25g}{70.906\frac{g}{mol}} = 0.35mols, \frac{73g}{32\frac{g}{mol}} = 2.28mols, \frac{50g}{28\frac{g}{mol}} = 1.79mols \tag{63}$$

$$0.35mols + 2.28mols + 1.79mols = 4.42mols \tag{64}$$

$$\frac{.35mols}{4.42mols} = 0.08, \frac{2.28mols}{4.42mols} = 0.52, \frac{1.79mols}{4.42mols} = .40 \tag{65}$$

$$Cl_{2PA} = 0.08, O_{2PA} = 0.52, N_{2PA} = 0.40 \tag{66}$$

10. Which gas will effuse faster, NH_3 or HCl? Explain why.

$$NH_3 \text{ because it has a lower molecular weight.} \tag{67}$$

11. Methane gas, CH_4, effuses through a barrier at a rate of 0.568 mL/min. If an unknown gas effuses through the same barrier at a rate of 0.343 mL/min, what is the molar mass of the gas?

$$\frac{\text{Rate A}}{\text{Rate B}} = \sqrt{\frac{\text{Molar Mass B}}{\text{Molar Mass A}}} \tag{68}$$

$$\frac{0.568\frac{mL}{min}}{0.343\frac{mL}{min}} = \sqrt{\frac{MM_B}{16\frac{g}{mol}}} \tag{69}$$

$$16\frac{g}{mol}(\frac{0.568\frac{mL}{min}}{0.343\frac{mL}{min}})^2 = MM_B \tag{70}$$

$$MM_B = 43.89\frac{g}{mol} \tag{71}$$

13 Tidbits

1. Double the moles, double the volume.
2. O Kelvin is the lowest possible temperature in the Universe. There is no maximum temperature.
3. 1 mole of any ideal gas has a volume of 22.4L.
4. If P doubles, V will Halve ?
5. If P doubles, T will Double ?
6. If P doubles, n will Double ?
7. Gases consist of molecules that are in ***continuous and random*** motion.
8. Volume of gas molecules is ***negligible*** in comparison with the total volume in which the gas is contained.
9. ***Attractive*** and ***repulsive*** forces between gas molecules are negligible.
10. Average kinetic energy is ***constant*** as long as T remains unchanged; collisions are ***elastic***.
11. Average kinetic energy is proportional to the absolute K temperature.
12. Real gases ***DO NOT*** follow the law perfectly because real particles have ***Attractive forces*** and ***Repulsive forces***. This is due to ***High pressures*** and ***Low temperatures***