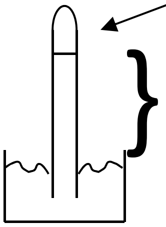
Gases

Chapter 5

Pressure

- Force per unit area
- Commonly measured in mm Hg
- Measured using a barometer



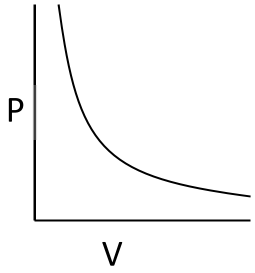
at sea level
= 760 mm Hg
=760 torr
= 1 atmosphere

(Evangelista Torricelli 1643)

The Gas Laws

Boyle's Law (Robert Boyle 1627-1691)

- The volume of a given amount of gas at constant temperature is inversely proportional to its pressure



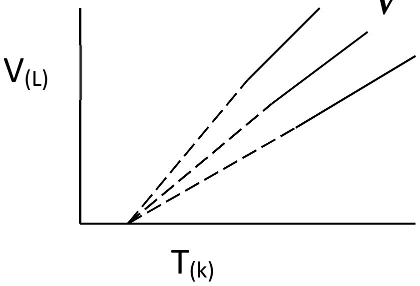
$$V \propto \frac{1}{P}$$

(5.2 p192)

Charles' Law (Jacques Charles 1746-1823)

- For a given mass of gas at constant pressure, volume is directly proportional to Kelvin temperature

$$V \propto T_k$$



$$\frac{V}{T} = k$$

$$\frac{V_1}{T} = \frac{V_2}{T}$$

(5.4 p194)

- Volumes of all gases extrapolate to zero at -273.2°C=absolute 0=0 K

Avogadro's Law

- At the same temperature and pressure, equal volumes of gases contain the same number of molecules
- At constant temperature and pressure, volume is directly proportional to the number of moles (n) of gas present

$$\frac{V}{n} = k \qquad \therefore \frac{V_1}{n_1} = \frac{V_2}{n_2} \text{ at constant T and P}$$

Ex. 5.5 (p195) 12.2L containing 0.50 mol O_2 @ 1 atm & 25°C

$O_2 \rightarrow$ ozone

Volume of ozone = ?

- Since: $V \propto \frac{1}{P}$, $V \propto T$, & $V \propto n$

  then $V \propto \frac{nT}{P}$ or $PV \propto nT$
- To make a proportion an equality, you need a proportionality constant

R = the universal gas constant

$$R = 0.0821 \frac{L \cdot atm}{mol \cdot K}$$

The Ideal Gas Law (or Equation of State)

$$PV = nRT$$

Ex. 5.6 (p197) $H_{2(g)}$ 8.56L at 0°C and 1.5atm

n = ?

Ex. 5.10 (p200) 0.35 mol Ar at 13°C and 568 torr
- if temperature is increased to 56°C and pressure is increased to 897 torr

$\Delta V = ?$

Molar Volume of a Gas at STP

$$V = \frac{nRT}{P} = \frac{(1.00 mol)(.0821 \frac{L \cdot atm}{mol \cdot K})(273K)}{1.00 atm}$$

$$V =$$

But only at STP!

5.13 (p203) 2.80L CH_4 at 25°C and 1.65atm and 35.0L of O_2 at 31°C and 1.25atm are ignited to form CO_2 at 125°C and 2.50atm. H_2O is also formed. What is the volume of the CO_2 at 125°C and 2.50atm?

A derivation (p204)

$$\text{n} = \frac{\text{grams of gas}}{\text{molar mass}} = \frac{\text{m}}{\text{molar mass}}$$

$$\text{PV} = \text{nRT}$$

$$\text{so: P} = \frac{\text{nRT}}{\text{V}} = \frac{(\frac{\text{m}}{\text{molar mass}})\text{RT}}{\text{V}} = \frac{\text{mRT}}{\text{(molar mass)V}}$$

$$\text{but: D} = \frac{\text{m}}{\text{V}}$$

$$\text{so: P} = \frac{\text{mRT}}{\text{V(molar mass)}} = \frac{\text{DRT}}{\text{molar mass}}$$

$$\text{so: molar mass} = \frac{\text{DRT}}{\text{P}}$$

Molar Mass of a Gas

$$\text{molarmass} = \frac{DRT}{P}$$

(This is NOT on the formula sheets!)

Ex. 5.14 (p204) at 1.50atm and 27°C D=1.95 g/L
mol wt. of gas = ?

Dalton's Law of Partial Pressure (John Dalton 1803)

- The total pressure exerted by a mixture of gases is the sum of the partial pressures of the individual gases.
- Partial pressure- the pressure a gas would exert if it were alone in a container

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

- "Does the identity of the gas make any difference in terms of pressure?"

$$PV=nRT$$

$$P=\frac{nRT}{V}$$ (no reference to molecular weight or gas indentity)

- The total number of moles of a gas in a container determines the pressure
- Gas identity is unimportant
- At the same conditions of temperature and pressure, 1mole of a gas exerts the same pressure as one mole of any other gas

Ex. 5.15 (p205) 46L O_2 at 25°C and 1.0 atm
12L He at 25°C and 1.0 atm } into a 5.0L tank
Calculate the partial pressure of each gas and the total pressure in the tank at 25°C.

Mole Fraction

- A concentration unit
- The ratio of the number of moles of one component in a mixture to the total number of moles in the mixture
- χ (chi) is the symbol

$$\chi_1 = \frac{n_1}{n_{total}}$$

Ex: 3mol O_2 & 1 mol N_2

$$\left.\begin{aligned} \chi_{O_2} &= \frac{3}{4} = .75 \\ \chi_{N_2} &= \frac{1}{4} = \underline{.25} \end{aligned}\right\} \chi \text{ does not have a unit}$$

1.0 } total sum of mole fractions must be 1.00!

- The relationship between gas pressures is the same as the mole fractions

Derivation:

$$PV = nRT$$

$$n_1 = \frac{P_1 V}{RT}$$

$$n_{total} = \frac{P_{total} V}{RT}$$

$$\chi_1 = \frac{n_1}{n_{total}} = \frac{\frac{P_1 V}{RT}}{\frac{P_{total} V}{RT}} = \frac{P_1\left(\frac{V}{RT}\right)}{P_{total}\left(\frac{V}{RT}\right)} = \frac{P_1}{P_{total}}$$

$$\boxed{\chi_1 = \frac{P_1}{P_{total}} = \frac{n_1}{n_{total}}}$$

Ex. 5.16 (p207) P_{O2}= 156 torr when atmospheric pressure is 743 torr.

χ_{O2}= ?

Ex. 5.17 (p208) In air χ_{N2}= 0.7808
If P_{total}= 760. torr
P_{N2}= ?

- When a gas is collected over H_2O, the sample of gas also contains water vapor.
- Gas Collection Apparatus:

- The higher the temperature, the greater the amount of $H_2O_{(g)}$ present.

Ex. 5.18 (p209) (Read the question)
$2KClO_{3\,(s)} \rightarrow 2KCl_{(s)} + 3O_{2\,(g)}$
P_{total}= 754torr, T= 22°C, at 21°C P_{H2O}=21torr
What mass of $KClO_{3\,(s)}$ was decomposed?

Kinetic Molecular Theory of Gases

- a series of statements used to explain the behavior of gases

Gases consist of particles that have the following properties:

1. the volume of the particles is so small compared to the distance between them that the volume of the particles is assumed to be zero
2. - the particles are in constant motion
   - pressure is caused by particles colliding with the container walls
3. the particles exert no force on each other (attraction or repulsion)
4. the average kinetic energy of the particles is directly proportional to the Kelvin temperature of the gas

- this describes an "Ideal Gas" (nonexistent)
- real gases behave most like the Ideal Gas at low pressure & high temperature
- real gases deviate from the Ideal Gas behavior because:

1. - individual particles do have volume
   - the volume available to nay given particle is less than the volume of the container
2. - real gas particles do attract each other
   - causes particles to hit container walls with slightly less force
   - the pressure exerted by the particles is lowered

Van der Waals Equation

(Johannes, 1873. 1910 Nobel Prize)

- corrects flaws in Kinetic Theory
- for real gases:

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

***** on formula sheet

P_{obs}= observed pressure

$\frac{n}{V}$ = concentration of particles

$$\left.\begin{array}{l} a = \\ b = \end{array}\right\}$$ proportionality constant determined from observing actual

experiments with a [particular gas

$$a = \frac{atm \cdot L^2}{mol^2} \qquad b = \frac{L}{mol}$$

*At normal conditions of temperature and pressure, real gases behave ideally enough to use PV=nRT

Root Means Square Velocity

- the average speed of a gas particle at a specific temperature

$U_{rms} = \sqrt{\overline{U}^2}$ $\qquad \overline{U}^2$ =the average of the squares of the particles velocities

$$\boxed{U_{rms} = \sqrt{\frac{3RT}{M}}}$$

* on formula sheet

R= constant = 8.31 (48) $\frac{J}{K \cdot mol}$
T= temperature in Kelvin
M= mass of one mole of particles in kilograms

Ex. 5.19 for He gas at 25°C, U_{rms}=?

Diffusion

- the rate at which gases mix with each other
- collisions between particles slow the diffusion
- the average distance traveled by a gas particle between collisions is called the mean free path
- rate of diffusion is inversely proportional to a particles mass

$$R \propto \frac{1}{\sqrt{m}}$$

$$\boxed{\frac{R_1}{R_2} = \frac{\sqrt{m_2}}{\sqrt{m_1}}}$$

- Also true for the effusion of a gas (passage through a small opening into an evacuated chamber)

Ex. 5.20 (p219) Calculate the ratio of the effusion rates of the gases hydrogen and uranium hexafluoride.