Chapter 10 - Gases

Section 10.1 - read for understanding
Section 10.2 - Pressure

$$\text{Pressure} = \frac{\text{Force}}{A} = \frac{\text{Newton}}{m^2} = \text{Pascal (Pa)}$$

Standard Atmospheric pressure is defined as the typical air pressure at sea level. It is equal to:

1 atm = 101.3 kPa = 760 mm Hg = 760 Torr

Although the SI unit of pressure is the Pascal, pressure is often described in any of the above units.

Converting of units of pressure in Sample Exercise 10.1 on p. 397.

a) 0.357 atm to torr

b) 6.6 x 10^{-2} torr to atm

c) 147.2 kPa to torr

Using a Manometer to Measure Gas Pressure – show students educator.com video

A sample of gas is placed in a flask attached to an open-end manometer. The level of mercury in the open end arm of the manometer has a height of 136.4 mm, and the mercury in the arm that is in contact with the gas has a height of 103.8 mm. What is the pressure of the gas in kPa, if the atmospheric pressure is 764.7 torr.

Now try questions 10.23 and 10.24 in your text.

Section 10.3 The Gas Laws (part of section 10.4 is included in here)

A. Boyle's law

- The volume of a fixed quantity of gas (maintained at constant temperature) is inversely proportional to the pressure. (ie when one gets larger the other gets smaller)

$$P_1 \times V_1 = P_2 \times V_2$$

Sample Problem:
A balloon is filled with 30. L of helium gas at 100. kPa. What is the volume when the balloon rises to an altitude where the pressure is only 25 kPa?

{note: you get a linear line when you plot V vs 1/p so it shows that V and P are inversely proportional to one another}

{note: breathing illustrates Boyle's Law: When the rib cage expands and the diaphragm moves downward the larger volume causes less pressure inside the lungs. Air pressure would then force air into the lungs}

B. Charles's Law

- The volume of a fixed amount of gas (at a constant pressure) is directly proportional to its Kelvin temperature.

  ** All gas laws require the Kelvin scale because negative numbers would cause negative answers for volume, moles etc.

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

  Sample Problem:
  A balloon inflated in an air-conditioned room at 27^{0}C, has a volume of 4.0 L. It is heated to a temperature of 57^{0}C. What is the new volume of the balloon if the pressure remains constant?

C. Avogadro's law

- The volume of a gas (maintained at constant temperature and pressure) is directly proportional to the number of moles of the gas.Equal volumes of different gases (at the same temperature and pressure) contain equal number of particles.

$$\frac{V_1}{n_1} = \frac{V_2}{n_2}$$

(ie 1 mole = 22.4 L for any gas at STP)

{These problems are just molar volume of a gas problems from grade 11 – no need to practice}

D. **Gay Lussac's Law**

- The pressure of a gas (maintained at constant volume) is directly proportional to its Kelvin temperature

$$\frac{P_1}{T_1} = \frac{P_2}{T_2}$$

Sample exercise:
A gas left in a used aerosol can is at a pressure of 100 kPa at 27^{0}C. If this can is thrown onto a fire, what is the internal pressure of the gas when its temperature reaches 927^{0}C?

Extra Practice Problems: Sample Exercise 10.5 and Practice Exercise on p. 405

Section 10.4 Combined Gas Law

The **Combined Gas Law** is: $\frac{P_1 \times V_1}{T_1} = \frac{P_2 \times V_2}{T_2}$ (the number of moles is fixed)

Sample Exercise 10.6 p. 406 (also try the practice problem at home)
An inflated balloon has a volume of 6.0 L at sea level (1 atm) and is allowed to ascend to an altitude until the pressure is 0.45 atm. During the ascent the temperature of the gas falls from 22^{0}C to -21^{0}C. Calculate the volume of the balloon at its final altitude.

Sections 10.4 and 10.5 Ideal Gas Law and Further Applications of the Ideal Gas Law

The **Ideal Gas Equation** is: $PV = nRT$ where R = the gas constant

- the value of R depends on the units of P, V, and T
- use values in Table 10.2 p. 402
- we normally use 8.314 J/mol- K or 0.08206 L-atm/mol-K

Remember that STP = standard temperature (0^0C) and pressure (1 atm or 101.3 kPa). Many properties of gases are tabulated at these conditions.

Note: you need to know the properties of real vs ideal gases from grade 11. {Show figure 10.13 on p. 403}

Ideal Gas Law Examples

a) A 8.5 L tank contains Helium gas at 16.9 atm and 25^0C. How many moles of He are available for making balloons?

b) If all the Helium was used to fill 1.5 L red balloons at 0.99 atm and 32^0C, how many balloons will you end up with?

Further Applications of the Ideal-Gas Equation

Gas Densities and Molar Mass

The Ideal gas equation can allow us to calculate gas density from molar mass.

$$PV = nRT$$

$$d = \frac{P \times MM}{RT} \quad \text{and} \quad MM = \frac{dRT}{P}$$

Determining Molar Mass of a Gas

a) An unknown gas is collected in an 850.0ml vessel at 1.02 atm and 23^0C. The evacuated vessel has a mass of 138.45 grams and the vessel and gas have a combined mass of 140.12 grams. Find the molar mass of the gas.

b) What is the density of carbon tetrachloride vapour at 714 torr and 125^0C?

c) An unknown gaseous hydrocarbon has a density of 3.24 g/L at 1.00 atm and 28^0C. Find its molar mass.

Another Application of the Ideal Gas Law

a) What volume of $CO_{2\,(g)}$ can be produced when 68 grams of propane, C_3H_8, burns in air at 25^0C and 1.02 atm?

$$C_3H_{8\,(g)} + 5\,O_{2\,(g)} \rightarrow 3\,CO_{2\,(g)} + 4\,H_2O_{\,(l)}$$

b) *Sample Exercise 10.9 p. 409* **Relating Volume of a Gas to the Amount of Another Substance in the Reaction**

The safety air bags in cars are inflated by nitrogen gas generated by the rapid decomposition of sodium azide, NaN_3.

$$2\,NaN_{3(s)} \rightarrow 2\,Na_{(s)} + 3\,N_{2(g)}$$

If an air bag has a volume of 36 L and is to be filled with nitrogen gas at a pressure of 1.15 atm at a temperature of 26.0^0C, how many grams of NaN_3 must be decomposed?

Section 10.6 Gas Mixtures and Partial Pressures

DALTON'S LAW of PARTIAL PRESSURES

- The total pressure of a mixture of gases equals the sum of the pressures that each would exert if it were present alone
- The pressure exerted by each gas is called the PARTIAL PRESSURE of that gas
- $P_{total} = P_1 + P_2 + P_3 + \ldots$

At a constant temperature and volume the total pressure is determined by the number of moles of gas present.

$$P_{total} = (n_1 + n_2 + n_3 \ldots)\frac{RT}{V} \quad \text{OR} \quad P_{total} = n_{total}\frac{RT}{V}$$

Sample Exercise 10.10 p. 411 **Applying Dalton's Law of Partial Pressures**
A gaseous mixture made from 6.00 grams of O_2 and 9.00 grams of CH_4 is placed in a 15.0 L vessel at 0^0C. What is the partial pressure of each gas, and what is the total pressure in the vessel?

A Side Topic – Concentration

Calculating Mole Fractions

Mole Fraction – the percent composition by moles of a single component in a mixture, represented in its decimal form

$$X_A = \frac{\text{Moles of one component } (n_A) \text{ in a mixture}}{\text{Sum of the moles of all components in the mixture}}$$

$$X_A = \frac{n_A}{n_A + n_B + n_C + n_D + \cdots + n_Z}$$

Ex1) Mole Fraction

Ex1) Find the mole fraction of each component in a gaseous solution that contains:

6.70 mol He, 2.50 mol Ar, and 1.60 mol Cl_2.

Ex2) Mole Fraction

Ex2) A gaseous solution contains Ne, H_2, and He. The mole fractions of Ne and H_2 are know to be 0.233 and 0.478 respectively. What is the mole fraction of He?

<u>Partial Pressure and Mole Fraction</u>

The ratio $\frac{n_1}{n_{tot}}$ is called the MOLE FRACTION of gas 1 which is called X_1

MOLE FRACTION, X, expresses the ratio of the number of moles of one component of a gas mixture to the total number of moles in the mixture.

$$P_1 = X_1 \bullet P_{total}$$

Sample Exercise 10.11 **Relating Mole Fractions and Partial Pressures** *p. 412*
A study of the effects of certain gases on plant growth requires a synthetic atmosphere composed of 1.5 mol percent CO_2, 18.0 mol percent O_2, and 80.5 mol percent Ar. ***(a)*** *Calculate the partial pressure of O_2 in the mixture if the total pressure of the atmosphere is to be 745 torr.* ***(b)*** *If this atmosphere is to be held in a 121-L space at 295 K, how many moles of O_2 are needed?*

Partial Pressure Using ICE charts

Ex2) The following reaction goes to completion in a 6.3 L vessel at 78°C.

$$\mathbf{H_{2(g)} + Cl_{2(g)} \rightarrow 2\ HCl_{(g)}}$$

The vessel contained 2.50 mol H_2 and 1.00 mol Cl_2 before the reaction took place.

a) Find the mole fraction of all gases present after the reaction is complete.
b) Find the total pressure in the vessel after the reaction.
c) Find the partial pressure of H_2 after the reaction.

Ex2) X_A and P_A (cont.)

a) Step 1. Find moles of each gas remaining after the reaction

	H_2 +	Cl_2 →	2 HCl
Initial			
Change			
Final			

Collecting Gas Over Water

For convenience in the laboratory, gases are often collected by bubbling them through water. This process causes the collected gas to be 'contaminated' with water vapour. That is, the collected gas and the water vapour both exert their own partial pressures.

$$P_{total} = P_{gas} + P_{water}$$

**The pressure exerted by water vapour at various temperatures is listed in Appendix B

Sample Exercise 10.12 ***Calculating the Amount of Gas Collected Over Water p.413***

A sample of $KClO_3$ is partially decomposed producing O_2gas that is collected over water. The volume of gas collected is 0.250 L at 26.0^{0}C and 765 Torr total pressure. **(a)** How many moles of oxygen are collected? **(b)** How many grams of $KClO_3$ were decomposed? *Answer: a) 0.00992 mol* O_2 *b) 0.810 g*

$$2\ KClO_{3(s)} \rightarrow 2\ KCl_{(s)} + 3\ O_{2(g)}$$

Section 10.7 Kinetic Molecular Theory

The KINETIC MOLECULAR THEORY is a model that explains the behavior of ideal gas particles. It includes:

1. All gases are made of particles moving in random motion
2. The volume of a gas is negligible compared to the total volume in which the gas is contained
3. Attractive forces between gas particles is negligible
4. Energy can be transferred between gas particles during collisions, but the average KE does not change over time as long as the temperature remains constant. In other words the collisions are ELASTIC.
5. The average KE of the particles is proportional to the gas's absolute temperature

Section 10.8 Molecular Effusion and Diffusion (continued…other notes in a power point)

ROOT - MEAN - SQUARE (rms) SPEED, *u*, =

- the speed of a molecule of a gas (not quite the same as average speed)
- refer to page 415

You can calculate the rms speed, *u*, of a gas by:

$$u = \frac{3RT}{MM}$$

Sample Exercise 10.14 **Calculating a Root-mean-Square Speed** p. 417
Calculate the rms speed, u, of a N_2 molecule at $25^{0}C$.

EFFUSION = the escape of gas particles through tiny holes

GRAHAM'S LAW OF EFFUSION = the effusion rate of a gas is inversely proportional to the square root of its molar mass (ie the smaller the molar mass of a gas the more rapidly it effuses)

The faster the molecule, the faster the rate of effusion. The smaller the molar mass of the molecule the faster the rate of effusion.

$$\frac{r_1}{r_2} = \frac{MM_2}{MM_1}$$

DIFFUSION =

- the spread of one substance throughout a space
- faster for lower molar mass molecules (same as effusion)
- collisions make diffusion more complicated that effusion. Collisions change the direction in which a particle is moving, therefore, the diffusion of a molecule from one point to another is made of many short (straight line) collision segments. The average distance travelled by a molecule between collisions is called the MEAN FREE PATH of the molecule.

Sample Exercise 10.15 **Applying Graham's Law** p. 419
An unknown gas is composed of homonuclear diatomic molecules diffuses at a rate that is only 0.355 times that of O_2 at the same temperature. Calculate the molar mass of the unknown and identify it.

Section 10. 9 Real Gases: Deviations from Ideal Behaviour

{Note: AP guide says that this only has to be understood qualitatively so we will not be using the van der Waals equation to calculate a numerical value for the deviation}

At high pressures (low volume) and low temperatures real gases deviate from ideal gas behavior. For an ideal gas we assume that the volume and attractive forces are negligible, but at high pressures and low temperatures these assumptions start to break down. The deviation is so great that the ideal gas equation needs to be adjusted to take these deviations into account. Under normal conditions the volume of a particle of gas can be ignored. Instead the volume of the space is used. But in small volumes (caused by high pressure) the volume of the gas particle does make a difference and needs to be taken into account.

The larger the molecule and increasing mass will cause a greater deviation from ideal behaviour, so Xe would deviate from ideal behavior the most (out of the noble gases).

The **van der Waals equation** is an equation that modifies the ideal gas equation to account for deviations:

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

The values for a and b are constants called correction values. They are different for each gas and they need to be looked up in data tables.