

## Chapter 13 – Gases

### 1) Sec 13.8 – Kinetic Theory of Gases and the Nature of Gases

The Kinetic Theory of Matter says that the tiny particles in all forms of matter are in constant motion. The energy that an object has because of its motion is kinetic energy.

#### Basic assumptions of the kinetic theory of gases

- Gases always spread out to fill the shape and volume of whatever space is available, so the volume of a gas is always the same as the volume of its container.
- Gases are composed of tiny particles (atoms or molecules) whose volume is negligible compared to the volume of the container they occupy.
- The particles in a gas move rapidly (average speed of air particles at room temperature is about 1000miles/hour) in constant random motion.
- The collisions that take place between gas particles with others or the walls of the container are said to be “completely elastic” (nothing is lost or gained).

### 2) Sec 13.1 – Gas Pressure

- Pressure is a result of the collisions of the gas particles with the sides of its container and is the force exerted per unit area (e.g., lbs/in<sup>2</sup>).
- Vacuum – The resulting empty space when no gas particles are present.
- Existence of Air Pressure — soda cans, balloon in a bottle, can.
- Measurement of Atmospheric Pressure – The barometer measures air pressure by determining the height of a column of mercury (d=13.6 g/cc) which can be supported by the air pressure.
- Units of pressure (mm Hg, in Hg, atmospheres, lbs/in<sup>2</sup>, kilopascals, torr)  
 $1 \text{ atm} = 760 \text{ mm Hg} = 760 \text{ torr} = 29.9 \text{ in Hg} = 14.7 \text{ lbs/in}^2 = 101.3 \text{ kPa}$
- When you work with gases, standard conditions (STP) are defined as a temperature of 0°C and a pressure of 760 mm Hg, 1 atm or 101.3 kPa.

#### Examples

1) Convert 190 mm Hg to atmospheres

2) Convert 190 mm Hg to KPa

Either:

Or:

3) Convert 720 mm Hg to lbs/in<sup>2</sup>

3) Sec 13.9 – Kinetic Theory and Kelvin Temperature

- a) The average kinetic energy of the particles of a substance is proportional to the absolute or Kelvin temperature of the substance ( $K = ^\circ C + 273$ ).
- b) The higher the temperature, the higher the average kinetic energy of the particles.
- c) The temperature at which the motion of particles ceases is known as absolute zero. ( $-273^\circ C$ ,  $-459^\circ F$ ,  $0 K$ )

4) The Effect of Adding or Removing a Gas

- a) If the number of gas molecules is doubled and the temperature and volume of the gas remain constant, the pressure will also double. That is, the pressure exerted by the gas is directly proportional to the amount of gas in the container. This is logical if one remembers that pressure is the result of the collisions between the gas particles and the walls of the container (think:  $lb/in^2$  – force per unit area). If one doubles the number of collisions, the pressure should also double.
- b) If the number of gas molecules is doubled and the volume is free to change, the pressure of the gas will be the same as the outside pressure. The volume will double, assuming that the temperature and pressure of the gas did not change.

5) Sec 13.2 – Boyle's Law for Pressure-Volume Changes

Reducing the volume of the container causes an increase in the gas pressure, and increasing the volume of the container causes a decrease in the gas pressure.

Boyle's Law – At constant temperature, the volume of a given mass of gas varies inversely with the pressure.  $P \times V = \text{constant}$  or  $P_1V_1 = P_2V_2$

Example: If a gas has a volume of 10 liters at a pressure of 2.0 atm, what must the volume change to in order for the pressure to be 6.5 atm?

6) Sec 13.3 – Charles' Law for Temperature-Volume Changes

Raising the temperature of a gas in a container either causes an increase in the pressure (if the volume of the container is fixed – cannot expand) or causes an increase in the volume of the container (if the container is free to expand). Since the average kinetic energy increases in proportion to the increase in the Kelvin temperature, the increase in either the pressure or the volume is directly proportional to the increase in the Kelvin temperature.

Charles' Law – At constant pressure, the volume of a given mass of gas varies directly with the absolute or Kelvin temperature.  $V \propto T$

$$\frac{V}{T} = \text{constant} \qquad \frac{V_1}{T_1} = \frac{V_2}{T_2}$$

Example: If a balloon has a volume of 10 Liters at 22°C, to what volume will it expand to at 86°C?

7) Gay-Lussac's Law for Temperature-Pressure Changes (not in book)

At constant volume, the pressure exerted by a given mass of gas varies directly with the absolute or Kelvin temperature.  $P \propto T$

$$\frac{P}{T} = \text{constant} \qquad \frac{P_{1-}}{T_1} = \frac{P_{2-}}{T_2}$$

Example: If the pressure of acetylene in a steel cylinder is 200 lbs/in<sup>2</sup> at a temperature of 0°C, what will be the pressure at 37°C?

8) The Combined Gas Law (not in book as a separate section, see page 320) – The three gas laws just discussed can be combined into a single expression called the Combined Gas Law.

For a given mass of gas, the product of its pressure and volume divided by its Kelvin temperature is a constant.

$$\frac{P \times V}{T} = \text{constant} \qquad \frac{P_1 \times V_1}{T_1} = \frac{P_2 \times V_2}{T_2}$$

This relationship is often used when different scientists want to compare identical reactions run under different conditions of temperature and pressure. Most often, however, it is used to express the results of experiments in terms of uniform STP (Standard Temperature and Pressure) conditions.

Example: Calculate the volume of a gas at 0°C and 760 mm Hg, if the volume of the gas was 500 mL when the gas was collected at 22°C and 720 mm Hg.

9) Sec 13.4 – Avagadro's Hypothesis: Volume and Moles – Equal volumes of gases at the same temperature and pressure contain equal numbers of particles. For example, at STP 22.4 L of any gas contains 1 mole or  $6.02 \times 10^{23}$  particles of gas. Note that this relates to the earlier mention of the relationship between volume and the amount of a gas.

The simple relationship is  $\frac{V}{n} = \text{constant}$  and  $\frac{V_{1-}}{n_1} = \frac{V_{2-}}{n_2}$

- 10) Sec 13.5 – The Ideal Gas Law combines the four physical properties (pressure, volume, temperature, and moles) for gases into one equation.

If one combines the laws, the expression becomes  $\frac{PV}{nT} = \text{constant} = R$

Where R is the combined proportionality constant and is called the **universal gas constant**. If one then turns this around, we have the **ideal gas law** in its usual form:

$$PV = nRT$$

P = pressure	V = volume in liters	
n = moles of gas	T = temperature in °K	

R is truly a constant. Its numeric value depends upon the units of pressure that are used for the calculation. Thus,

$$R = 0.0821 \text{ L-atm/mole-K} = 8.31 \text{ L-KPa/mole-K} = 62.4 \text{ L-mm Hg/mole-K}$$

and one can convert between the various R values by converting between the pressure units.

Further, since  $n = \frac{m}{MW}$  Then  $PV = \frac{mRT}{MW}$       m = mass in grams  
 MW = molecular mass (g/mole)

If mass is given or needed, then use the second equation; if moles are given or needed, then use the first equation.

Examples:

- a) You fill a rigid steel cylinder with a volume of 20 L of nitrogen gas to a pressure of 20,000 KPa at 27°C. How many moles of nitrogen does the cylinder contain?
- b) Calculate the mass in grams of CO<sub>2</sub>, if 200 mL were collected over water at a temperature of 22°C and a total pressure of 755 mm Hg.
- c) What pressure will be exerted by 25.5 g of ammonia, NH<sub>3</sub>, gas at 25°C in a 5.0 L container?

10) Sec 13.6 – Dalton's Law of Partial Pressures – The total pressure exerted by a mixture of gases is the sum of their individual pressures.  $P_{\text{total}} = P_a + P_b + P_c + \text{etc.}$

a) The partial pressure of a gas in mixture of gases is the pressure it would exert on the container if it were all alone in the same volume.

Example: What is the partial pressure of the oxygen in the atmosphere if the atmospheric pressure is 760 mm Hg and 20% of the air molecules are oxygen?

b) Collecting a Gas over Water – In the laboratory, many gases are collected by water displacement. If a volume measurement of the gas is made, this volume includes not only the gas being considered, but also water vapor. It then becomes necessary to make the proper allowance for this water vapor which is done by subtracting the vapor pressure that water exerts at that temperature from the total pressure that has been measured.

$$P_{\text{gas}} = P_{\text{measured}} - \text{v.p. of water}$$

The vapor pressure of water varies with the temperature of the water. (See water vapor table.)

Example: Suppose you prepared a sample of nitrogen and collected it over water at 15°C at a total pressure of 745 mm Hg. This sample occupied 310 mL. What would be the volume of the dry gas at STP?

11) Molecular Mass of a Gas from Experimental Data

Since all gases have the same volume at a given temperature and pressure (for example 22.4L at STP), this fact can be used to determine the molecular mass of any substance that is a gas or can be changed into a gas.  $MW = \frac{mRT}{PV}$

Examples:

1) Calculate the molecular mass of a gas if 0.840 g of the gas occupied a volume of 505 mL at 27°C and 740 mm Hg.

2) a) Calculate the empirical formula of a compound that contains 52.2% carbon, 13.0% hydrogen, and 34.8% oxygen.

- b) Calculate the molecular mass of the compound if 0.371 g of the compound when vaporized occupied a volume of 250 mL at 100°C and 750 mm Hg.
- c) What is the molecular formula of the compound?

13) Sec 13.7 & 13.10 – Real Versus Ideal Gases

A gas that adheres to the gas laws at some conditions of temperature and pressure is said to exhibit ideal behavior. No gas behaves ideally, however, at all temperatures and pressures. The conditions at which gases behave ideally are at high temperatures (well above the boiling point of the substance) and at low pressure. Gases do not obey the gas laws at low temperatures (near the temperature when they are about to condense to a liquid) and at very high pressures (100 times atmospheric pressure).

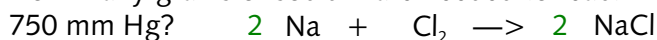
14) Sec 13.11 – Stoichiometry Problems Involving Gases not at STP

Use  $n = \frac{PV}{RT}$  to calculate the moles if the volume of a gas is given at other than STP.

Use  $V = \frac{nRT}{P}$  to calculate the volume of a gas at conditions other than STP.

Examples:

- 1) How many grams of sodium are needed to react with 400 mL of chlorine at 25°C and 750 mm Hg?

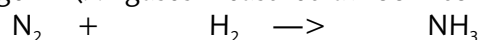


- 2) What volume of carbon dioxide at a pressure of 5.0 atm and a temperature of 127°C will be produced by the reaction of 53 g of sodium carbonate with excess hydrochloric acid?



- 15) Gay-Lussac's Law of Combining Volumes (not in book) – The volumes of different gases involved in a reaction, if measured at the same conditions, are in the same ratio as these coefficients for these gases in the balanced chemical equation.

Example: What volume of ammonia can be produced by the reaction of 200 mL of hydrogen with excess nitrogen? (All gases measured at room conditions.)



- 16) Weinkauff Modification of the Ideal Gas Law

When looking at the ideal gas law, one can also see how we could arrange it in proportions the way we did the other laws. The resulting statement would be:  $\frac{P_1 V_1}{n_1 T_1} = \frac{P_2 V_2}{n_2 T_2}$

If one remembers the ideal gas law itself and then this statement, one can have all the other laws readily and answer most gas questions.

- \* We have already talked about the ways the ideal gas law itself can be used.
- \* If the number of moles and temperature are constant, the statement just becomes Boyle's Law.
- \* If the number of moles and the pressure are constant, then we have Charles' Law.
- \* If the pressure and temperature are constant, then we have Avogadro's Law.