Unit 9 Chemistry

**Chapter 14: The Behavior of Gases**

1. **Properties of gases**:
   1. Gases are compressible: There is space between the particles.
   2. Gases have kinetic energy (measured by temperature) and move to take the shape and volume of their container.
   3. Four variables are used to describe a gas: Temperature (Kelvin), Pressure (Kilopascals), Volume (Liters) and Moles.

* o C + 273 = Kelvin
* 101.3 kpa = 1 atmosphere = 760 mm Hg = 29.8 inches of Hg
* 1000 ml = 1 liter
* 1 mole = \_\_\_\_\_grams (periodic table and round to .1 grams)
  1. Gases *change states* by changing Temperature, Pressure, and Volume.

A. A Gas is defined as:

1. **KE (Kinetic Energy):**

**Temperature**:

Rise in average KE of particles causes temperatures of substances to rise

\_\_\_\_\_\_Kinetic Energy = \_\_\_\_\_\_\_\_Temperature

Substance cools, particles move more slowly and KE decreases

\_\_\_\_\_\_Kinetic Energy = \_\_\_\_\_\_\_\_Temperature

Absolute zero:

Temperature conversions:

1. **Gas Pressure**

\_\_\_\_\_\_\_\_\_ in collisions causes an \_\_\_\_\_\_\_\_\_\_ in pressure

\_\_\_\_\_\_\_\_\_ in collisions causes a decrease in pressure

Empty space with no particles has no pressure = \_\_\_\_\_\_\_\_\_

How do you measure pressure? Use a \_\_\_\_\_\_\_\_\_\_\_\_

Pressure Conversion :

Example Problem: Convert 780 mm Hg into kilopascals.

1. **Gas Laws**: a mathematical relationship of the variables that effect gases.

These laws are based on an “ideal” gas. Ideal gases do not exist.

* 1. An **ideal gas** is a gas that remains a gas under all temperature, pressure, and volume changes. (This allows for linear graphs and equations to be developed.)
  2. **Real gases** become liquids and solids with changes in temperature, pressure, and volume.
  3. Therefore, gas laws do not work under extreme changes in temperature, pressure, and volume. They only work under the normal environmental ranges of change found on the earth.

1. The Pressure-Volume Relationship: **BOYLES LAW**
   1. At a constant temperature (Kelvin), the volume of a gas varies inversely or indirectly with pressure. (When one goes up, the other goes down: ↓P=↑V and ↑P =↓V )
   2. P1 x V1 = constant K

P2 x V2 = constant K Therefore **P1 x V1 = P2 x V2**

Ex: The pressure on 3.5 L of anesthetic gas changes from 105 kpa to 40.5 kpa. What will be the new volume of the gas if the temperature remains constant?

Ex: A gas with a volume of 4.0 L at a pressure of 205 kpa is allowed to expand to a volume of 12.000 ml. What is the pressure of the gas (in kpa) in the container if the temperature is constant?

Ex: A 2.5 L helium balloon at 1.15 atm. of pressure is released by a child in a parking lot. The balloon travels upward and the atmospheric pressure is now 87.0 kpa. What would be the new volume of the balloon? (Assume a constant temperature.)

1. The Temperature-Volume Relationship: **CHARLES’ LAW**
   1. At a constant pressure, the volume of a gas is directly proportional to its Kelvin temperature. (When one goes up, the other goes up: ↓T =↓V and ↑T =↑V)
   2. V1/T1 = constant

V2/T2 = constant Therefore: **V1 = V2**

**T1 T2**

Ex. If a sample of gas occupies 6.8 liters at 325 ºC, what will be its volume at 25 ºC if the pressure remains constant?

Ex. Exactly 5000 ml of air at -50.0 ºC is warmed to 100.0 ºC. What is the new volume in liters, if the pressure remains constant?

1. The Temperature-Pressure Relationship: **GAY-LUSSAC’S LAW**
   1. a. At a constant volume, the pressure of a gas is directly proportional to the Kelvin temperature. (When one goes up, the other goes up: ↓T =↓P and ↑T =↑P)
   2. P1/T1 = constant

P2/T2 = constant Therefore: **P1 = P2**

**T1 T2**

Ex. The gas left in an aerosol can is at a pressure of 103 kpa at 25 ºC. If this can (volume is constant) is thrown into a fire, what is the pressure on the gas inside the can if the temperature reaches 928 ºC.

1. The **COMBINED GAS LAW**
   1. Boyles’ plus Charles’ plus Gay-Lussac’s Law

b. **P1 V1 = V2 P2**

**T1 T2**

Ex. A gas at 155 kpa and 25 ºC occupies a container with an initial volume of 1.00 L. By changing the volume, the pressure of the gas increases to 6400 mm Hg as the temperature goes to 125 ºC. What is the new volume of the gas?

Ex. A 5000 ml air sample at -50 ºC and 107 kpa is heated to 102 ºC and expands to 7.00 L. What is the new pressure of the gas in kpa?

1. **IDEAL GAS LAW**
   1. **Ideal gas constant: R** and how it is derived.

P1 V1 = V2 P2 Multiply both sides by 1

T1 T2n

P1 V1 = V2 P2  If P2 = 101.3 kpa, V2 = 22.4 L, T2 = 273 K, n = 1 mole

T1 n1 T2 n then the left side of the equation = **8.31 L kpa/ K moles = R**

P1 V1 = R Therefore P1 V1 = n1 R T1 or **PV = n RT**

T1 n1

* 1. Use the ideal gas law if moles = n or grams are mentioned. It allows a comparison of all gases to 1 mole of gas at STP and brings in the mass component.

n = moles x Xgrams = grams or grams x 1 mole = moles = n

1 mole Xgrams

Ex. A child has a lung capacity of 2.20 L. How many grams of air (molar mass of air = 29.0 grams/1 mole) do her lungs hold at a pressure of 102 kpa and a normal body temperature of 37 ºC?

Ex. A deep underground cavern contains 2.24 x 106 liters of methane (CH4) at a pressure of 1.50 x 103 kpa and a temperature of 42 ºC. How many grams of CH4 does this natural gas deposit contain?

Ex. .386 moles of gas is collected at 30.0 ºC and 730 mm Hg. What volume cylinder will the gas occupy?

Ex. A chemist collect 5900 ml of SO2 when the atmospheric pressure is .950 atmospheres and the mass of the SO2 collected is 80.0 grams. At what temperature, in ºC, did the chemist collect the gas?

1. Gas Molecules: Mixtures and Movements
   1. **DALTON’S LAW:** At a constant volume and temperature, the total pressure exerted by a mixture of gases is equal to the sum of the partial pressure of the component gases.

Ex. Air contains O2, N2, CO2, and trace amounts of other gases. What is the partial pressure of the O2 at 101.3 kpa if the partial pressure of N2 = 79.10 kpa, CO2 = .040 kpa, and trace amounts = .94 kpa?

1. **GRAHAM’S LAW**
   1. Diffusion: the movement of particles from areas of higher concentrations to areas of lower concentrations.
   2. Effusion: the process by which a gas escapes through a small hole in a container.
   3. Graham’s Law

**Rate A (light gas) = √molar mass B (heavy gas)**

**Rate B (heavy gas) = √molar mass A (light gas)**

Note: Heavy particles escape slower than lighter particles according to the equation above.

**Lighter gas A escapes \_\_\_\_ times faster that the heavier gas B.**

Ex. Compare the rates of effusion for a He and N2 leaky tire.

Ex. Compare the rates of effusion for a CH4 and SO2.