PROPERTIES OF GASES

- ***Gases*** are the ***least dense*** and ***most mobile*** of the three phases of matter.

- ***Particles of matter*** in the gas phase are ***spaced far apart*** from one another and ***move rapidly*** and ***collide*** with each other often.

- Gases occupy much greater space than the same amount of liquid or solid. This is because the gas particles are spaced apart from one another and are therefore compressible. Solid or liquid particles are spaced much closer and cannot be compressed further.

- Gases are characterized by four properties. These are:

    1. Pressure (P).
    2. Volume (V)
    3. Temperature (T)
    4. Amount (n)

Kinetic-Molecular Theory of Gases

- Scientists use the ***kinetic-molecular theory (KMT)*** to describe the ***behavior of gases***. The KMT consists of several postulates:

    1. Gases consist of ***small particles*** (atoms or molecules) that move ***randomly*** with ***rapid velocities***.

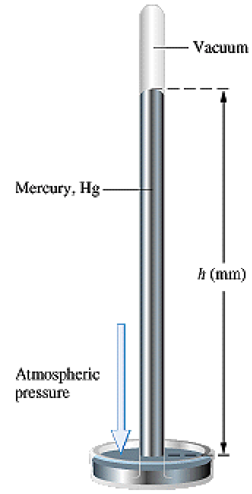
    2. Gas particles have ***little attraction*** for one another. Therefore, ***attractive forces*** between gas molecules can be ***ignored***.

    3. The ***distance*** between the particles is ***large*** compared to their size. Therefore the ***volume*** occupied by gas molecules is ***small*** compared to the volume of the gas.

    4. Gas particles move in ***straight lines*** and ***collide*** with each other and the container ***frequently***. The force of ***collisions*** of the gas particles with the walls of the container causes pressure.

    5. The ***average kinetic energy*** of gas molecules is ***directly proportional*** to the ***absolute temperature*** (***Kelvin***).

PRESSURE AND ITS MEASUREMENT



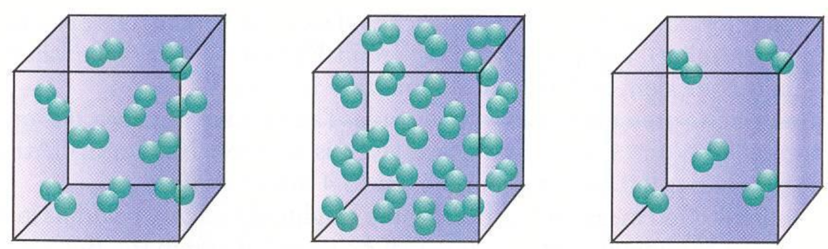
- ***Pressure*** is the result of ***collision*** of gas particles with the sides of the container. Pressure is defined as the ***force per unit area***.

- ***Pressure*** is measured in ***units of atmosphere*** (***atm***) or ***mmHg*** or ***torr***. The SI unit of pressure is pascal (Pa) or kilopascal (kPa).

1 atm = 760 mmHg
1 mmHg = 1 torr
1 atm = 101.325 kPa

- Atmospheric ***pressure*** can be ***measured*** with the use of a ***barometer***. Mercury is used in a barometer due to its high density. At sea level, the mercury stands at 760 mm above its base.

- The ***pressure*** of a gas is ***directly proportional*** to the ***number of particles*** (***moles***) present.



1 mol H_2 $P = 1$ atm | 2 mol H_2 $P = 2$ atm | 0.5 mol H_2 $P = 0.5$ atm

<u>Examples:</u>

1. The atmospheric pressure at Walnut, CA is 740. mmHg. Calculate this pressure in torr and atm.

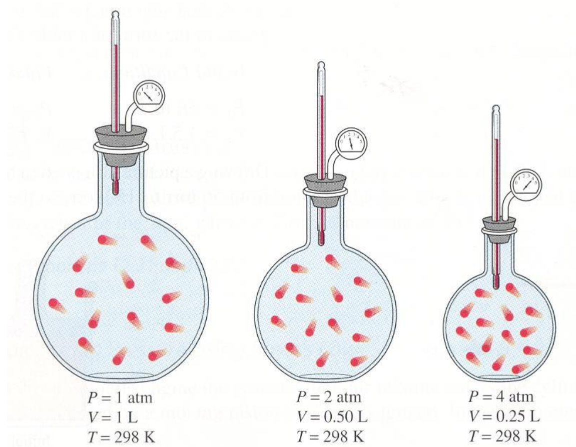
   1 mmHg = 1 torr therefore, 740. mmHg =

   740. mmHg x ———————— = atm

2. The barometer at a location reads 1.12 atm. Calculate the pressure in mmHg and torr.

RELATIONSHIP BETWEEN PRESSURE & VOLUME
BOYLE'S LAW

- At ***constant temperature***, the ***volume of a fixed amount*** of gas is ***inversely proportional*** to its ***pressure***.



$$\mathbf{P_1V_1 = P_2V_2}$$

Examples:

1. A sample of H_2 gas has a volume of 5.0 L and a pressure of 1.0 atm. What is the new pressure if the volume is decreased to 2.0 L at constant temperature?

   $P_1 = 1.0$ atm $\quad P_2 = ???$ $\qquad P_2 = \dfrac{P_1V_1}{V_2} = \dfrac{(1.0\text{ atm})(5.0\text{ L})}{2.0\text{ L}} = 2.5$ atm

   $V_1 = 5.0$ L $\quad V_2 = 2.0$ L

2. A sample of gas has a volume of 12-L and a pressure of 4500 mmHg. What is the volume of the gas when the pressure is reduced to 750 mmHg?

   $P_1 =$ $\qquad P_2 =$

   $V_1 =$ $\qquad V_2 = ???$

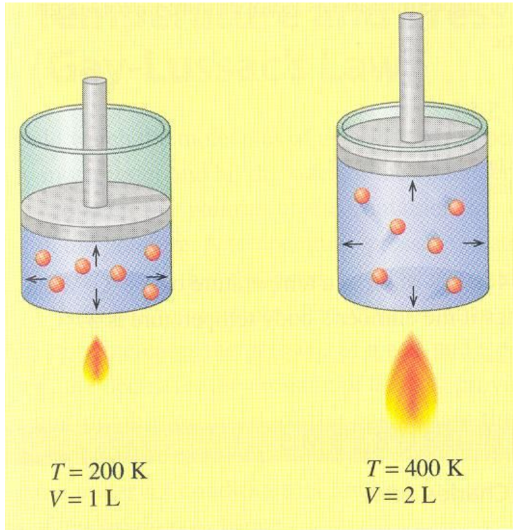
3. A sample of hydrogen gas occupies 4.0 L at 650 mmHg. What volume would it occupy at 2.0 atm?

RELATIONSHIP BETWEEN TEMP. & VOLUME
CHARLES'S LAW

- At ***constant pressure***, the ***volume of a fixed amount*** of gas is ***directly proportional*** to its ***absolute temperature.***

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

Note: T must be in unit of K



Examples:

1. A 2.0-L sample of a gas is cooled from 298K to 278 K, at constant pressure. What is the new volume of the gas?

$V_1 = 2.0$ L $\quad V_2 = ???$ $\quad V_2 = V_1 \times \dfrac{T_2}{T_1} = 2.0 \text{ L} \times \dfrac{278 \text{ K}}{298 \text{ K}} = 1.9 \text{ L}$

$T_1 = 298$ K $\quad T_2 = 278$ K

2. A sample of gas has a volume of 5.0-L and a temperature of 20°C. What is the volume of the gas when the temperature is increased to 50°C, at constant pressure?

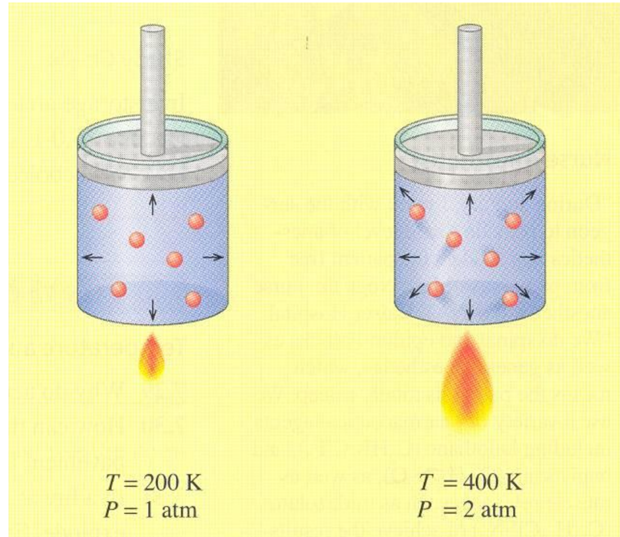
$V_1 =$ $\quad V_2 =$

$T_1 =$ $\quad T_2 =$

3. If 20.0 L of oxygen gas is cooled from 100 °C to 0 °C, what is the new volume?

RELATIONSHIP BETWEEN TEMP. & PRESSURE
GAY-LUSSAC'S LAW

- At ***constant volume***, the ***pressure of a fixed amount*** of gas is ***directly proportional*** to its ***absolute temperature.***



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

Examples:

1. An aerosol spray can has a pressure of 4.0 atm at 25°C. What pressure will the can have if it is placed in a fire and reaches temperature of 400°C ?

P_1 = 4.0 atm P_2 = ??? $P_2 = P_1 \times \frac{T_2}{T_1} = 4.0 \text{ atm} \times \frac{673 \text{ K}}{298 \text{ K}} = 9.0 \text{ atm}$

T_1 = 298 K T_2 =673 K

2. A cylinder of gas with a volume of 15.0-L and a pressure of 965 mmHg is stored at a temperature of 55°C. To what temperature must the cylinder be cooled to reach a pressure of 850 mmHg?

P_1 = P_2 =

T_1 = T_2 =???

3. The pressure of a container of helium is 650 mmHg at 25 °C. If the container is cooled to 0 °C, what will the pressure be?

VAPOR PRESSURE & BOILING POINT

- In an open container, liquid molecules at the surface that possess sufficient energy, can break away from surface and become gas particles or vapor. In a closed container, these gas particles can accumulate and create pressure called ***vapor pressure***.

- ***Vapor pressure*** is defined as the pressure above a liquid at a given ***temperature***. Vapor pressure varies with each liquid and increases with temperature. Listed below is the vapor pressure of water at various temperatures.

Temperature (°C)	Vapor Pressure (mmHg)
0	5
10	9
20	18
30	32
37	47[a]
40	55
50	93
60	149
70	234
80	355
90	528
100	760

[a]At body temperature.

- A liquid reaches its boiling point when its vapor pressure becomes equal to the external pressure (atmospheric pressure). For example, at sea level, water reaches its boiling point at 100°C since its vapor pressure is 760 mmHg at this temperature.

- At higher altitudes, where atmospheric pressure is lower, water reaches boiling point at temperatures lower than 100°C. For example in Denver, where atmospheric pressure is 630 mmHg, water boils at 95°C, since its vapor pressure is 630 mmHg at this temperature.

RELATIONSHIP BETWEEN PRESSURE, VOL. & TEMP.
COMBINED GAS LAW

- All ***pressure-volume-temperature*** relationships can be combined into a single relationship called the ***combined gas law***. This law is useful for studying the effect of ***changes in two*** variables.

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

- The individual gas laws studied previously are embodied in the combined gas law.

$$\frac{P_1V_1}{\color{red}{T_1}} = \frac{P_2V_2}{\color{red}{T_2}}$$ **Boyle's Law**

$$\frac{\color{red}{P_1}V_1}{T_1} = \frac{\color{red}{P_2}V_2}{T_2}$$ **Charles's Law**

$$\frac{P_1\color{red}{V_1}}{T_1} = \frac{P_2\color{red}{V_2}}{T_2}$$ **Gay-Lussac's Law**

Examples:

1. A 25.0-mL sample of gas has a pressure of 4.00 atm at a temperature of 10°C. What is the volume of the gas at a pressure of 1.00 atm and a temperature of 18°C?

$P_1 =$ $P_2 =$

$V_1 =$ $V_2 =$

$T_1 =$ $T_2 =$

$$V_2 = V_1 \text{ x } \frac{P_1}{P_2} \text{ x } \frac{T_2}{T_1}$$

$$V_2 = 25.0 \text{ mL x } \frac{4.00 \text{ atm}}{1.00 \text{ atm}} \text{ x } \frac{291 \text{ K}}{283 \text{ K}} = 103 \text{ mL}$$

2. A sample of ammonia has a volume of 20.0 mL at 5 °C and 700 mmHg. What is the volume of the gas at 50°C and 850 torr?

$P_1 =$ $P_2 =$

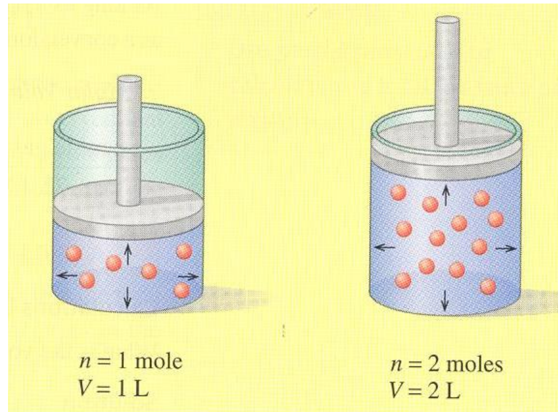
$V_1 =$ $V_2 =$

$T_1 =$ $T_2 =$

RELATIONSHIP BETWEEN VOLUME & MOLES
AVOGADRO'S LAW

- At ***constant temperature and pressure***, the ***volume of a fixed amount*** of gas is ***directly proportional*** to the ***number of moles.***

- As a result of Avogadro's Law, ***equal volumes*** of different gases at the ***same temperature and pressure*** contain ***equal number of moles*** (molecules).



- This relationship also allows chemists to relate volumes and moles of a gas in a chemical reaction. For example:

$$\mathbf{2\,H_2\,(g) \;+\; O_2\,(g) \;\rightarrow\; 2\,H_2O\,(g)}$$

2 molecules	1 molecule	2 molecules
2 moles	1 mole	2 moles
2 Liters	1 Liter	2 Liters

Examples:

1. A sample of helium gas with a mass of 18.0 g occupies 1.6 Liters at a particular temperature and pressure. What mass of oxygen would occupy1.6 L at the same temperature and pressure?

$$\text{mol of He} = 18.0\ \cancel{\text{g}}\ \text{x}\ \frac{1\ \text{mol}}{4.00\ \cancel{\text{g}}} = 4.50\ \text{mol He}$$

$$\text{mol of } O_2 = \text{mol He} = 4.50\ \text{mol} \qquad \text{(since at same T \& P)}$$

$$\text{mass of } O_2 = 4.50\ \cancel{\text{mol}}\ \text{x}\ \frac{32.0\ \text{g}}{1\ \cancel{\text{mol}}} = 144\ \text{g}$$

2. How many Liters of NH_3 can be produced from reaction of 1.8 L of H_2 with excess N_2, as shown below?

$$N_2\,(g) + 3\,H_2\,(g) \rightarrow 2\,NH_3\,(g)$$

$$1.8\ \cancel{\text{L H}_2}\ \text{x}\ \frac{2\ \text{L NH}_3}{3\ \cancel{\text{L H}_2}} = 1.2\ \text{L NH}_3$$

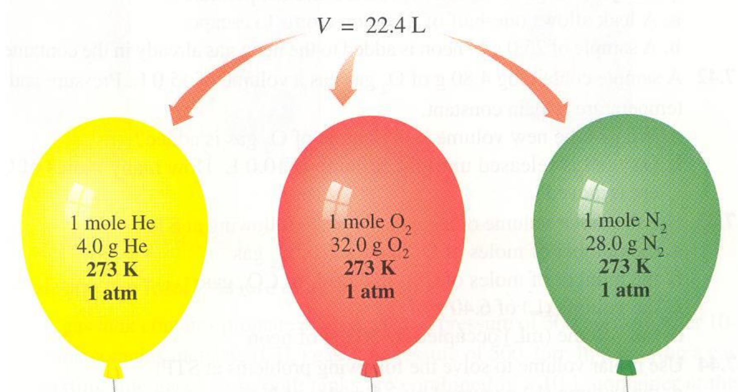
STP & MOLAR VOLUME

- To better understand the factors that affect gas behavior, a set of ***standard conditions*** have been chosen for use, and are referred to as ***Standard Temperature and Pressure (STP)***.

STP: 0°C (273 K) and 1 atm (760 mmHg)

- At ***STP*** conditions, ***one mole of any gas*** is observed to occupy a ***volume of 22.4 L***. This quantity is referred to as ***Molar Volume***.

Molar Volume of a gas at STP = 22.4 L



<u>Examples:</u>

1. If 2.00 L of a gas at STP has a mass of 3.23 g, what is the molar mass of the gas?

$$\text{mol of gas} = 2.00\ \cancel{\text{L}} \times \frac{1\ \text{mol}}{22.4\ \cancel{\text{L}}} = 0.0893\ \text{mol}$$

$$\text{molar mass} = \frac{\text{g}}{\text{mol}} = \frac{3.23\ \text{g}}{0.0893\ \text{mol}} = 36.2\ \text{g/mol}$$

2. A sample of gas has a volume of 2.50 L at 730 mmHg and 20°C. What is the volume of this gas at STP?

P_1 = P_2 =

V_1 = V_2 =

T_1 = T_2 =

IDEAL GAS LAW

- Combining all the laws that describe the behavior of gases, one can obtain a useful relationship that relates the volume of a gas to the temperature, pressure and number of moles.

$$\mathbf{V = \frac{nRT}{P}}$$

$$\mathbf{R = Universal\ Gas\ Constant = 0.0821\ \frac{L\ atm}{mol\ K}}$$

- This relationship is called the ***Ideal Gas Law***, and commonly written as:

$$\mathbf{\underset{atm}{\underset{\uparrow}{P}}\ \underset{L}{\underset{\uparrow}{V}} = \underset{mol}{\underset{\uparrow}{n}}\ R\ \underset{K}{\underset{\uparrow}{T}}}$$

Examples:

1. A sample of H_2 gas has a volume of 8.56 L at a temperature of 0°C and pressure of 1.5 atm. Calculate the moles of gas present.

   P= 1.5 atm n= ??? $\mathbf{n = \frac{PV}{RT}}$

   V= 8.56 L T= 273 K $n = \frac{(1.5\text{ atm})(8.56\text{ L})}{(0.0821\text{ L atm/mol K})(273\text{ K})} = 0.57\text{ mol}$

2. What volume does 40.0 g of N_2 gas occupy at 10°C and 750 mmHg?

   P= n=

   V=??? T= $V = \frac{n\,RT}{P} =$

3. A 23.8-L cylinder of contains oxygen at 20.0 °C and 732 mmHg. How many moles of oxygen does it contain?

PARTIAL PRESSURES
DALTON'S LAW

- Many gas samples are mixture of gases. For example, the air we breathe is a mixture of mostly oxygen and nitrogen gases.

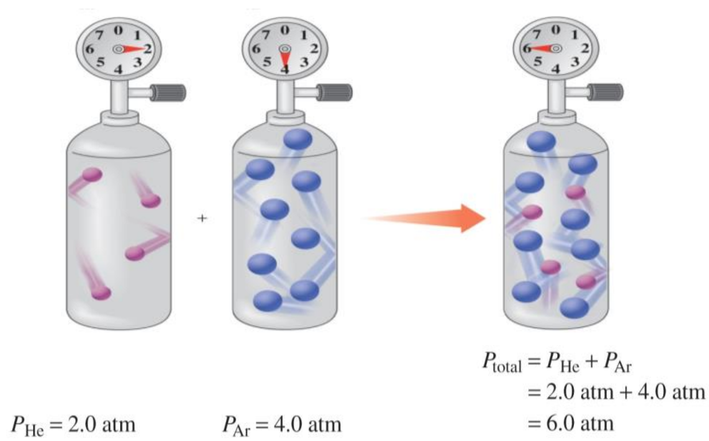
- Since gas particles have no attractions towards one another, each gas in a mixture behaves as if it is present by itself, and is not affected by the other gases present in the mixture.

- In a mixture, each gas exerts a pressure as if it was the only gas present in the container. This pressure is called ***partial pressure*** of the gas.

- In a mixture, the sum of all the partial pressures of gases in the mixture is equal to the total pressure of the gas mixture. This is called ***Dalton's law*** of partial pressures.

$$P_{total} = P_1 + P_2 + P_3 + \cdots$$

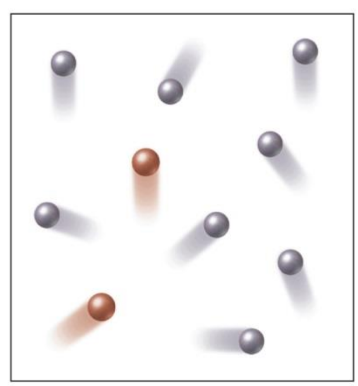
Total pressure of a gas mixture **=** ***sum of the partial pressures of the gases in the mixture***



- The partial pressure of each gas in a mixture is proportional to the amount (mol) of gas present in the mixture. Therefore, the partial pressure of a gas is its fractional composition (X) times the total pressure of the gas. (Note: fractional composition is percent divided by 100)

- For example, in the mixture shown below, since the mixture is 80% He and 20% Ne, the partial pressure of each gas can be calculated as shown below:

$$P_{He} = X_{He}\, P_{tot} = (0.80)(1.0 \text{ atm}) = 0.80 \text{ atm}$$
$$P_{Ne} = X_{Ne}\, P_{tot} = (0.20)(1.0 \text{ atm}) = 0.20 \text{ atm}$$



Gas mixture (80% He, 20% Ne)
$P_{tot} = 1.0$ atm
$P_{He} = 0.80$ atm
$P_{Ne} = 0.20$ atm

PARTIAL PRESSURES

Examples:

1. Two 10-L tanks, one containing propane gas at 300 torr and the other containing methane at 500 torr, are combined in a 10-L tank at the same temperature. What is the total pressure of the gas mixture?

$$P_{total} = P_1 + P_2 = 300 \text{ torr} + 500 \text{ torr} = 800 \text{ torr}$$

2. A scuba tank contains a mixture of oxygen and helium gases with total pressure of 7.00 atm. If the partial pressure of oxygen in the tank is 1140 mmHg, what is the partial pressure of helium in the tank?

$$P_{oxygen} \text{ (in atm)} = 1140 \text{ mmHg} \times \frac{\quad}{\quad} =$$

$$P_{total} = P_{oxygen} + P_{helium}$$

$$P_{helium} =$$

3. A mixture of gases contains 2.0 mol of O_2 gas and 4.0 mol of N_2 gas with total pressure of 3.0 atm. What is the partial pressure of each gas in the mixture?