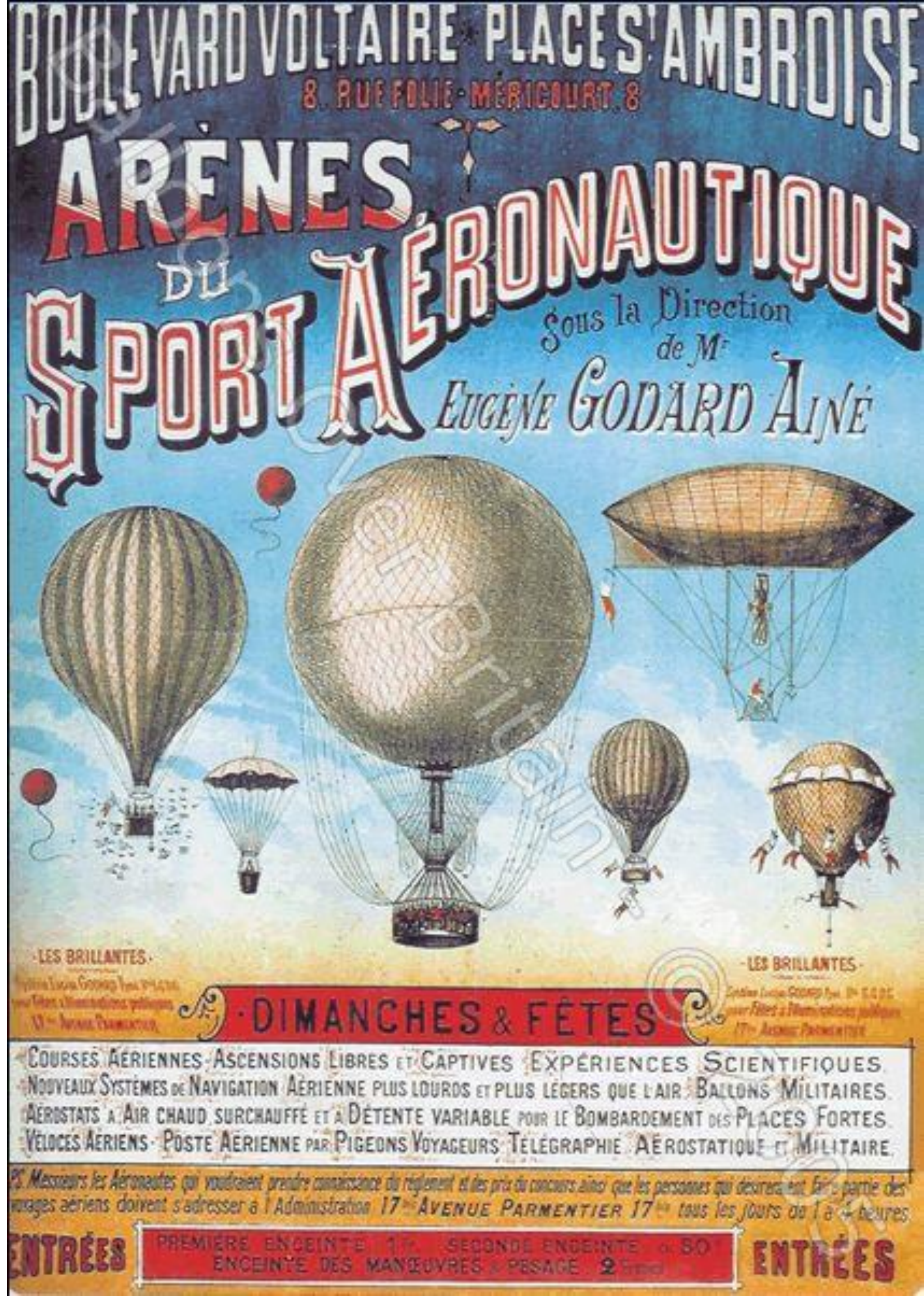
Gases and Gas Laws



Kinetic Molecular Theory

- Particles of matter are ALWAYS in motion

- **Volume of individual particles is ≈ zero. Consists of large number of particles that are very far apart**

- **Collisions of particles with container walls cause pressure exerted by gas but are negated**

- **Particles exert no forces on each other (neither attraction or repulsion)**

- **Average kinetic energy ∝ Kelvin temperature of a gas. *(the warmer the faster)***

The Meaning of Temperature

$$(\mathrm{KE})_{\mathrm{avg}} = \frac{3}{2} RT$$

- Kelvin temperature is an index of the random motions of gas particles

- **(higher *T* means greater motion.)**

Kinetic Energy of Gas Particles

At the same conditions of temperature, all gases have the same average kinetic energy. Therefore at the same temperature lighter gases are moving FASTER

$$KE = \frac{1}{2}mv^2$$

m = mass

v = velocity

Measuring Pressure

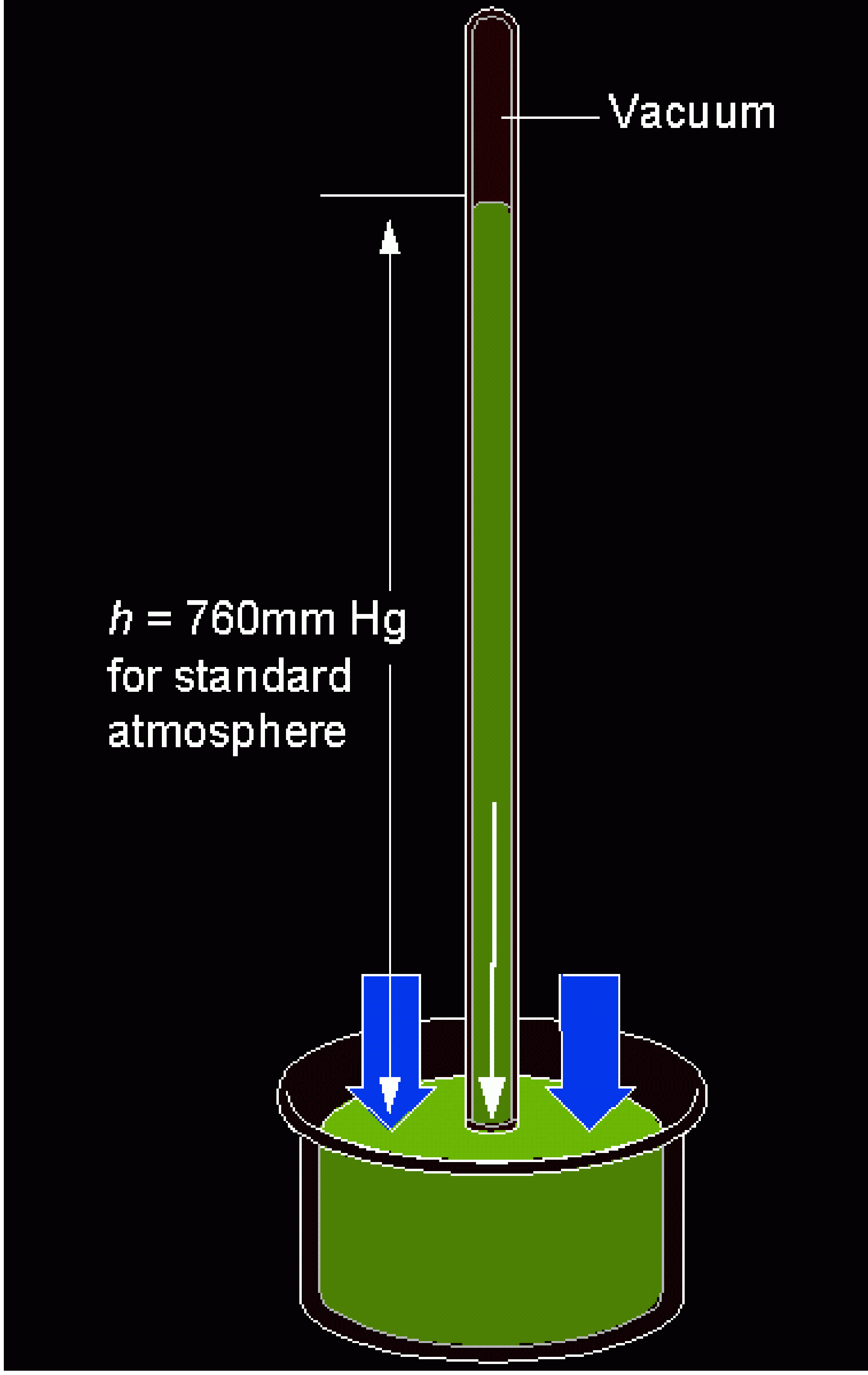
The first device for measuring atmospheric pressure was developed by Evangelista Torricelli during the 17th century.

The device was called a "barometer"

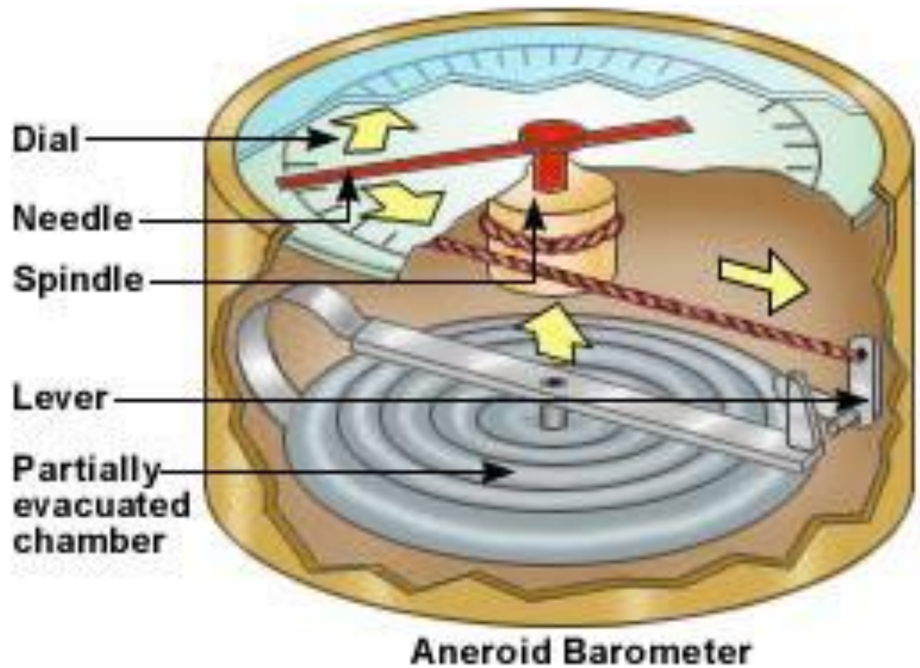
- Baro = weight
- Meter = measure

An Early Barometer

The normal pressure due to the atmosphere at sea level can support a column of mercury that is 760 mm high



The Aneroid Barometer



Pressure

- Is caused by the collisions of molecules with the walls of a container

- is equal to force/unit area

- SI units = Newton/meter^2 = 1 Pascal (Pa)

- 1 standard atmosphere = 101.3 kPa
- 1 standard atmosphere = 1 atm

- unit conversions:

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

Units of Pressure

Unit	*Symbol*	*Definition/Relationship*
Pascal	Pa	SI pressure unit 1 Pa = 1 newton/meter^2
Millimeter of mercury	mm Hg	Pressure that supports a 1 mm column of mercury in a barometer
Atmosphere	atm	Average atmospheric pressure at sea level and 0 °C
Torr	torr	1 torr = 1 mm Hg

Mathematical Computations:

- Convert the following:

- 737 mmHg to atm
- **97.5 kPa to atm**
- **0.650 atm to mmHg**

.785 atm to torr

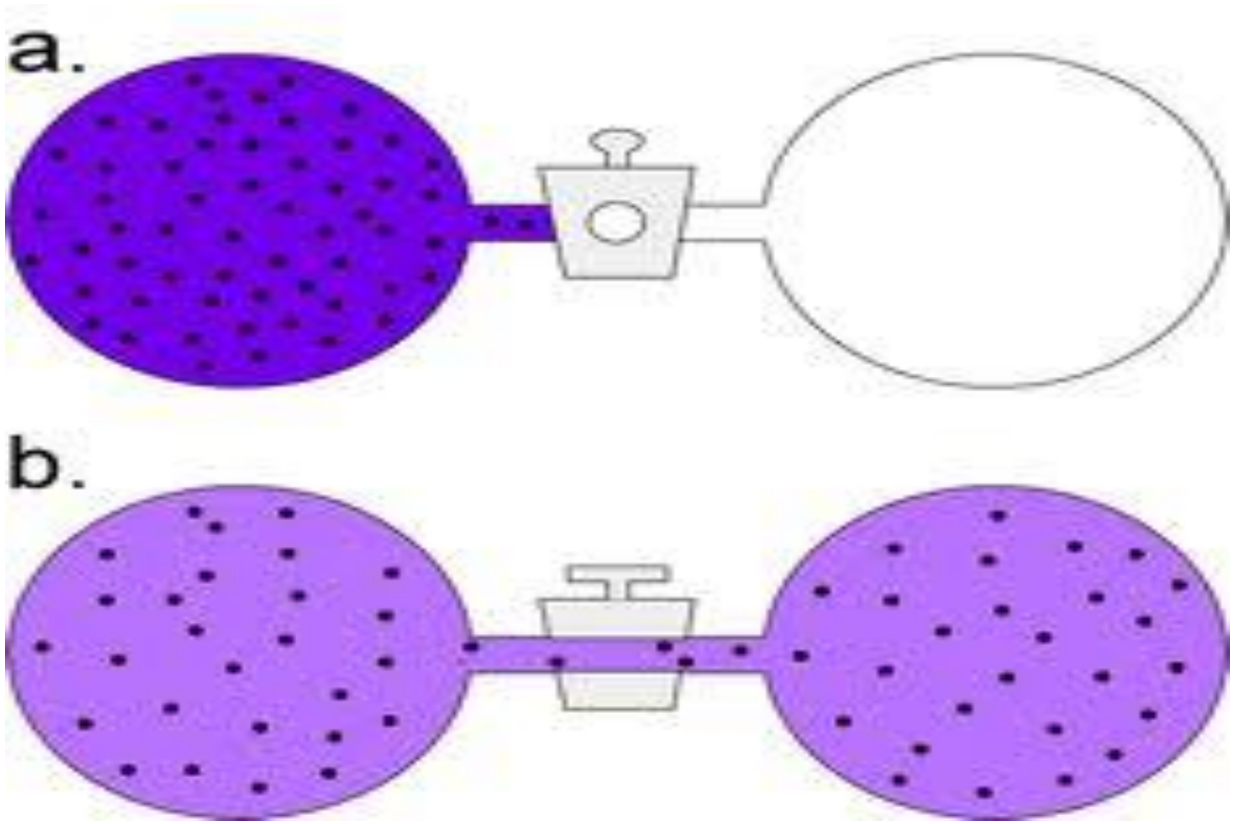
Physical Properties of Gases

1. Gases have low density

- **1/1000 the density of the equivalent liquid or solid**

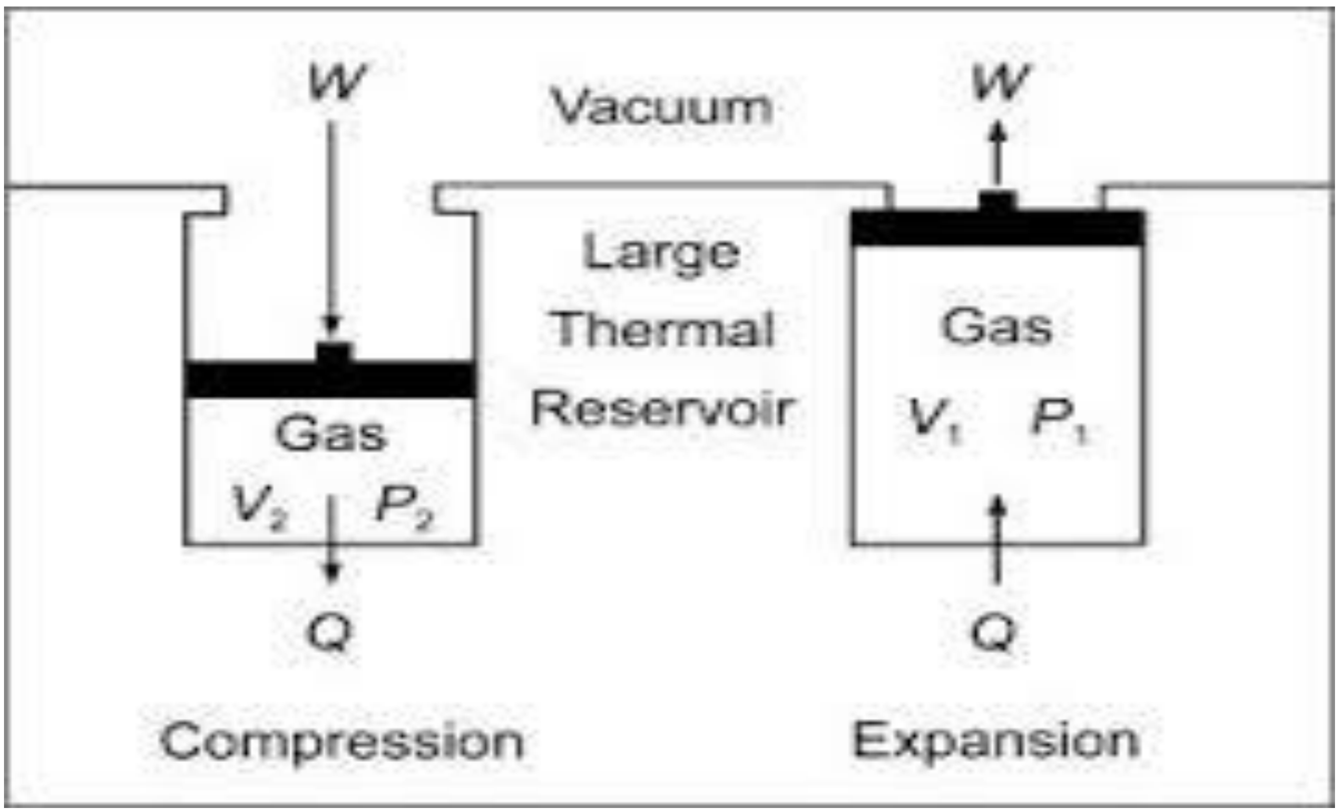
The Nature of Gases

- **2 Expansion-** particles move away and fill up the space available.



3. Compression

- Molecules can be moved together- so close a gas can be liquefied

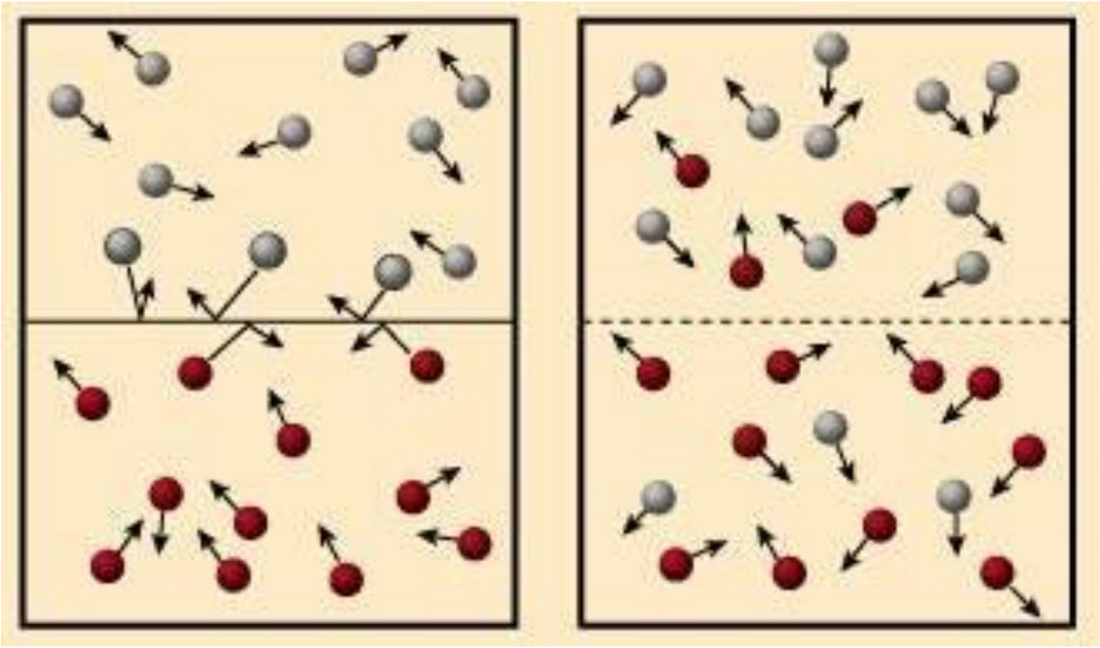


4. Fluidity

- Gases like liquids can be poured. Molecules slide past one another.

5. Diffusion

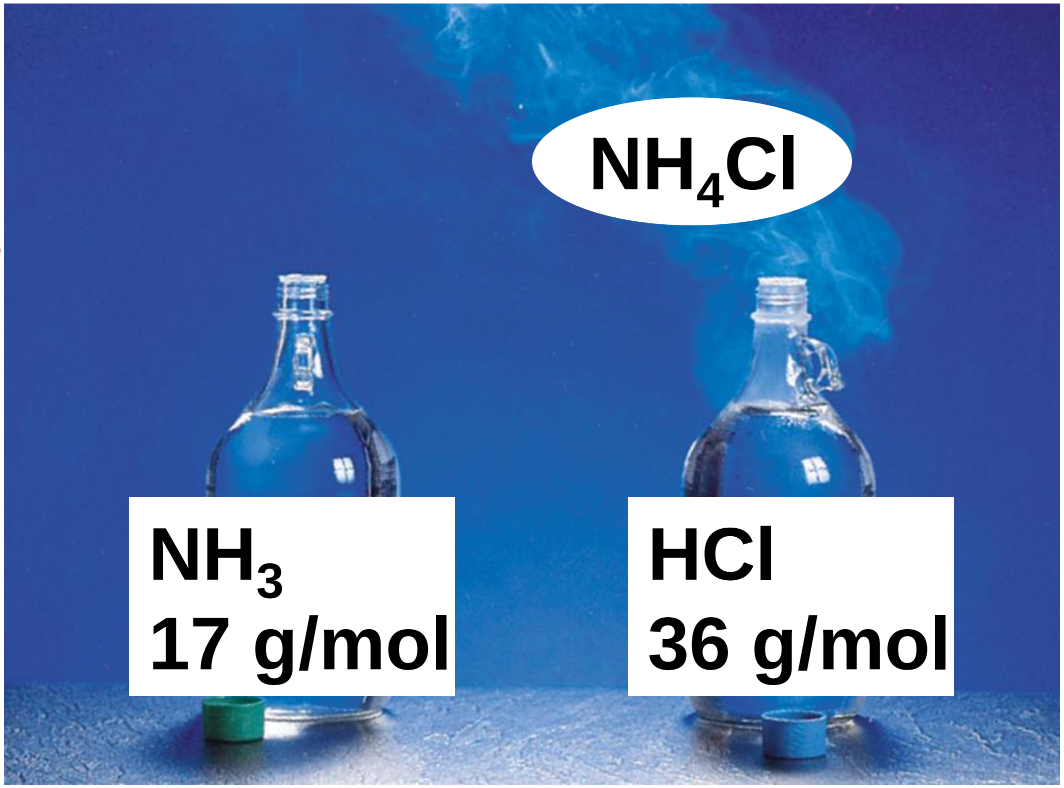
Diffusion: describes the mixing of gases. The rate of diffusion is the rate of gas mixing.



- Diffusion: *The spontaneous mixing of molecules caused by the movement of molecules in the liquids or gases-*
- *(The warmer the faster the diffusion)*

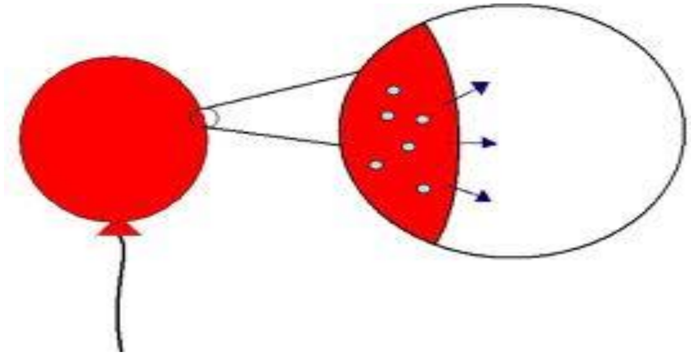
Gas diffusion- is the gradual mixing of molecules of one gas with molecules of another by virtue of their kinetic properties. (depends on <u>Temp</u>, Size, Shape, and polarity)

Consider the reaction of ammonia and hydrochloric acid given that both reagents are at the same temperature:



Effusion: gas is forced through a tiny opening

The balloon is porous and gas will slowly escape



Or even pumping up a tire

Standard Temperature & Pressure "STP"

- P = 1 atmosphere (atm),

  760 torr

  101.3 kPa

  760 mmHg

- $T = 0^\circ$ C, 273 Kelvins

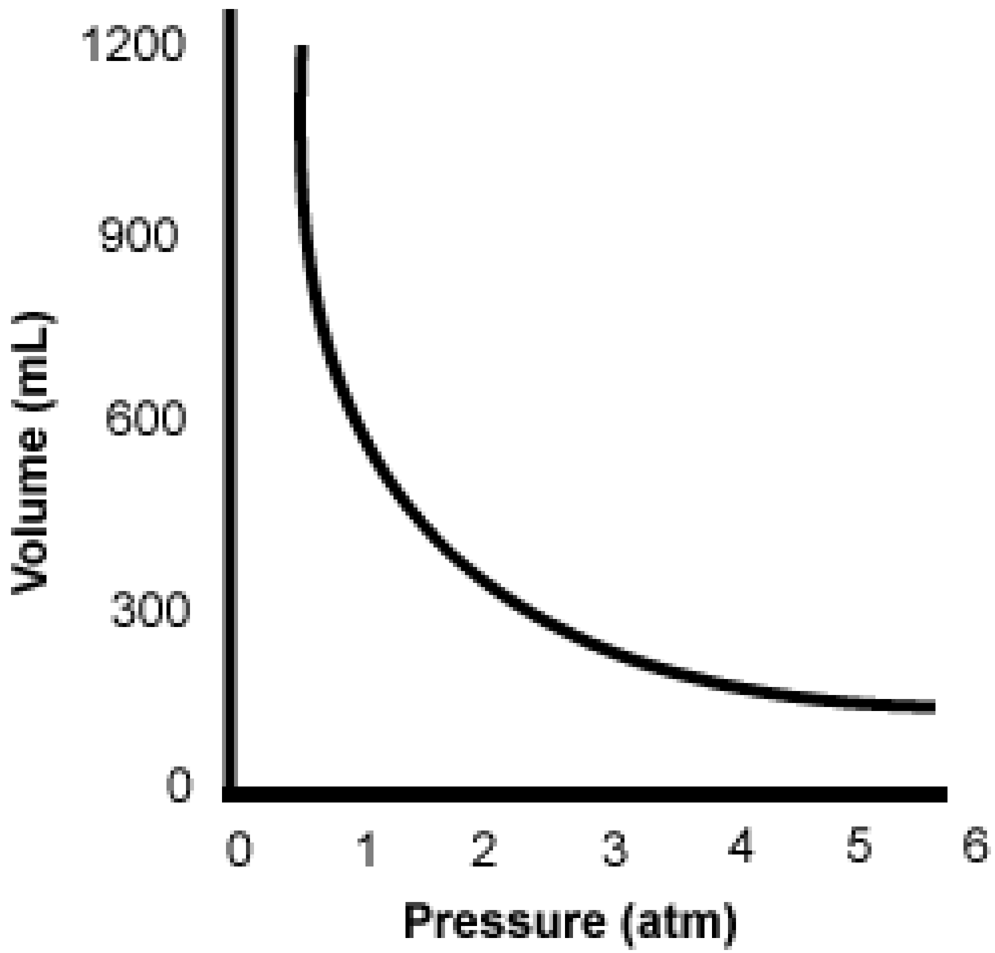
Boyle's Law

The volume of a fixed mass of gas varies inversely with the pressure at a constant temperature (pressure and volume under both conditions must have the same units)

Think of a piston in a plunger: it is harder to push the plunger down at the end of the stroke where volume has decreased)

$$P_1V_1 = P_2V_2$$

A Graph of Boyle's Law



Converting Celsius to Kelvin

Gas law problems involving temperature require that the temperature be in KELVINS!

$$\text{Kelvins} = {}^\circ C + 273.15$$

$$^\circ C = \text{Kelvins} - 273.15$$

Temperature Computations:

- Covert 22.0^{0}c to Kelvin

- **Convert 285 Kelvin to ^{0}c**

Charles's Law

- The volume of a gas at constant pressure is directly proportional to Kelvin temperature, and extrapolates to zero at zero Kelvin.

  (P = constant)

Think about an inflated balloon in a warm home. You step out into the cold night air. What happens?

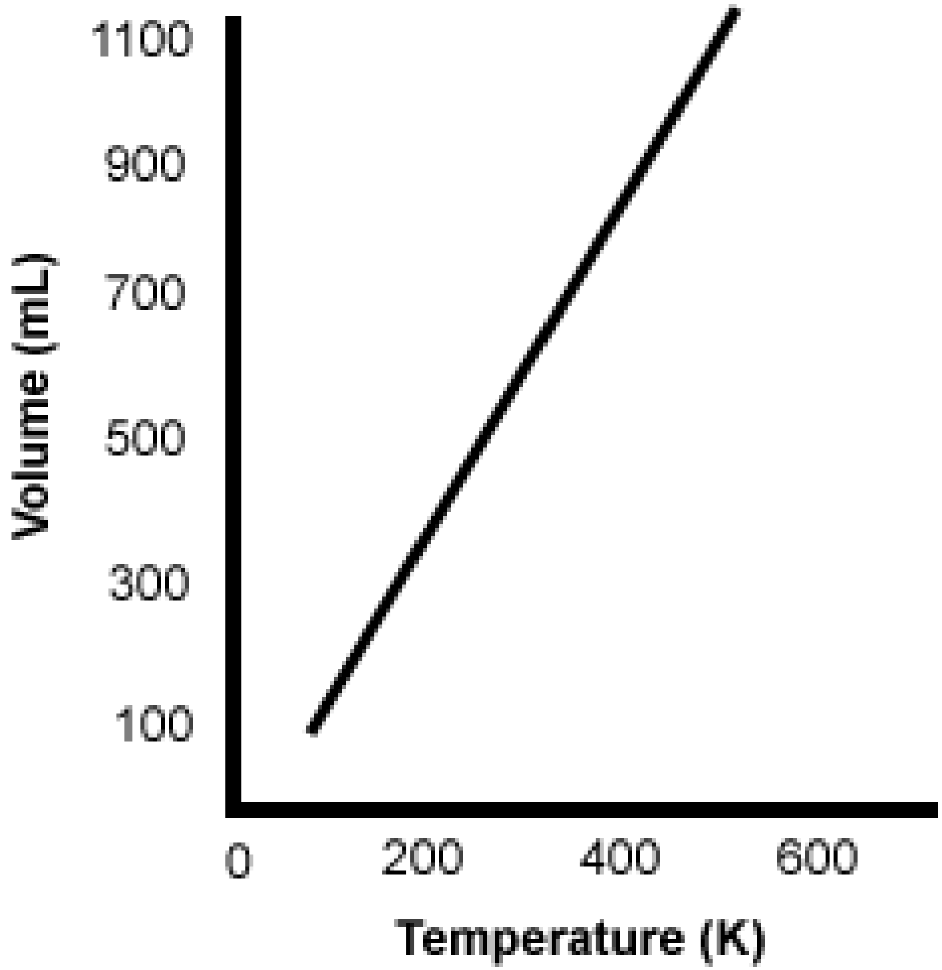
$$\frac{V_1}{T_1} = \frac{V_2}{T_2} \qquad (P = \text{constant})$$

Temperature MUST be in <u>KELVINS!</u>

Mathematical Computation : Charles Law #1

- A sample of neon in a balloon at constant pressure has a volume of 725.ml at 22.0^{0}c. If the balloon is left in a car where the volume expeands to 800.00 ml, what temperature in 0c is the interior of the car assuming the pressure remains constant?

A Graph of Charles' Law



Gay Lussac's Law

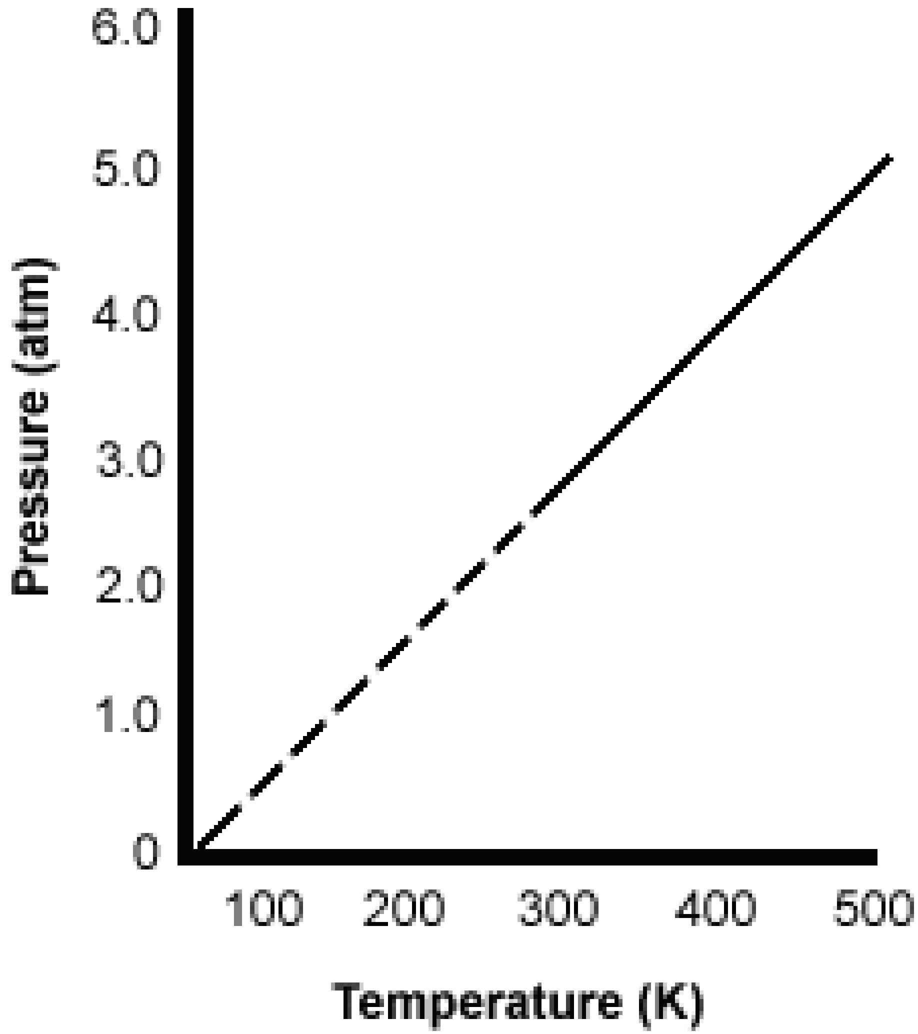
The pressure and Kelvin temperature of a gas are directly related, provided that the volume remains constant.

Think about when you have a fixed space and increase the temp, molecules will increase movement which increases collisions which is pressure (and vice versa)!

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

Temperature MUST be in <u>KELVINS</u>

A Graph of Gay-Lussac's Law



The Combined Gas Law

The combined gas law expresses the relationship between pressure, volume and temperature of a fixed amount of gas.

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

Boyle's law, Gay-Lussac's law, and Charles' law are all derived from this by holding a variable constant. Again temperature must be in Kelvins and units must match!

Mathematical Computation: Combined Gas Law #1

- A sample of oxygen gas occupies a volume of 0.250 liters at 22.0^0c and exerts a pressure of 750.0 mmHg. What volume will the gas occupy if the temperature drops to 5.0^0c and the pressure decreased to 0.950 atm?

Dalton's Law of Partial Pressures

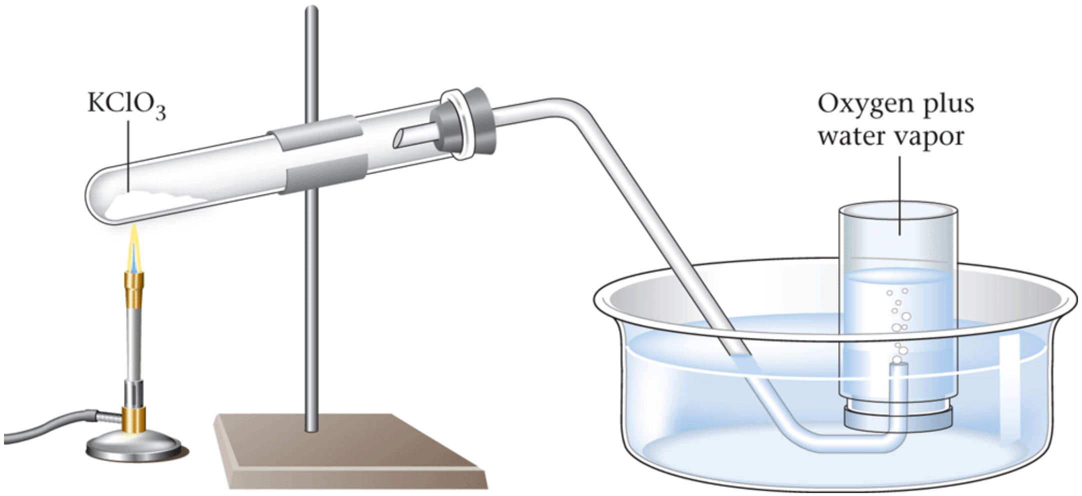
For a mixture of gases in a container, The total pressure is equal to the sum of the parts

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

(This is particularly useful in calculating the pressure of gases collected over water. Water vapor exerts a particular pressure at particular temperature –water vapor chart)

B. Dalton’s Law of Partial Pressures

Collecting a gas over water



- Total pressure is the pressure of the gas + the vapor pressure of the water.

Collecting Hydrogen

Vapor Pressure at various temperatures:

T ^{0}c	P mmHg	T ^{0}c	P mmHg	T ^{0}c	P mmHg
0	4.58	32	37.0	70	233.7
5	6.54	35	42.2	80	355.1
10	9.21	37	47.07	90	525.8
15	12.79	40	55.3	100	760
20	17.54	45	73.9		
22	19.8	50	92.5		
25	23.76	55	121		
27	26.7	60	149.4		
30	31.8	65	192		

Ideal Gases

Ideal gases are imaginary gases that perfectly fit all of the assumptions of the kinetic molecular theory.

- Gases consist of tiny particles that are far apart relative to their size.

- Collisions between gas particles and between particles and the walls of the container are elastic collisions

- No kinetic energy is lost in elastic collisions

Ideal Gases (continued)

- Gas particles are in constant, rapid motion. They therefore possess kinetic energy, the energy of motion
- There are no forces of attraction between gas particles
- The average kinetic energy of gas particles depends on temperature, not on the identity of the particle.

Real Gases Do Not Behave Ideally

Real gases **DO** experience inter-molecular attractions

Real gases **DO** have volume

Real gases **DO NOT** have elastic collisions

Deviations from Ideal Behavior

Likely to behave nearly ideally	***Likely not to behave ideally***
Conditions: Gases at high temperature and low pressure	Conditions: Gases at low temperature and high pressure
Charateristics: Small symmetrical non-polar gas molecules	Characteristics Large, non-symmetrical polar gas molecules

Ideal Gas & Gas Stoichiometry

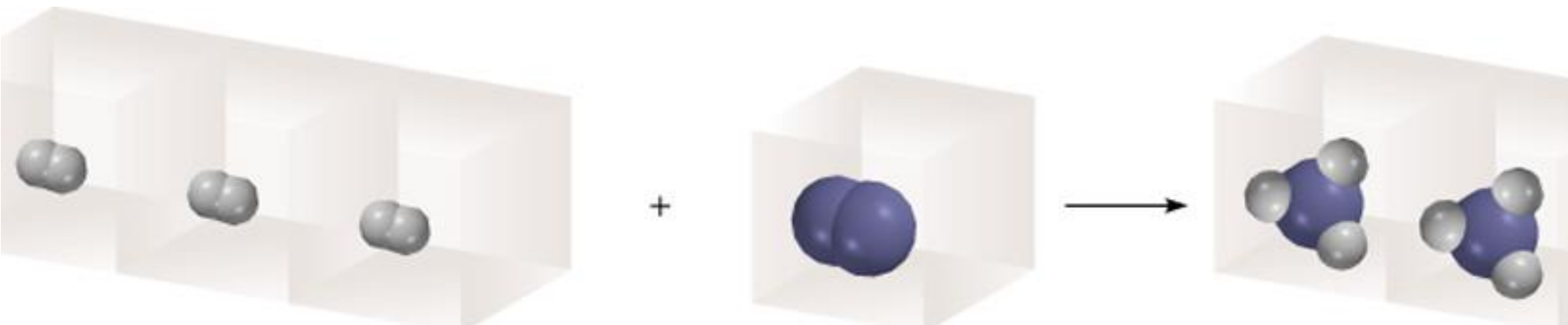
Avogadro's Law

$V \propto$ number of moles (n)

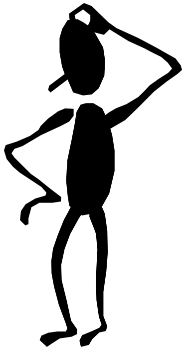
V = constant x n

$V_1/n_1 = V_2/n_2$

Constant temperature
Constant pressure



$3H_2\,(g)$	+	$N_2\,(g)$	$\longrightarrow$	$2NH_3\,(g)$
3 molecules	+	1 molecule	$\longrightarrow$	2 molecules
3 moles	+	1 mole	$\longrightarrow$	2 moles
3 volumes	+	1 volume	$\longrightarrow$	2 volumes



Ammonia burns in oxygen to form nitric oxide (NO) and water vapor. How many volumes of NO are obtained from one volume of ammonia at the same temperature and pressure?

$4 \; NH_3 + 5 \; O_2 \longrightarrow 4NO + 6H_2O$

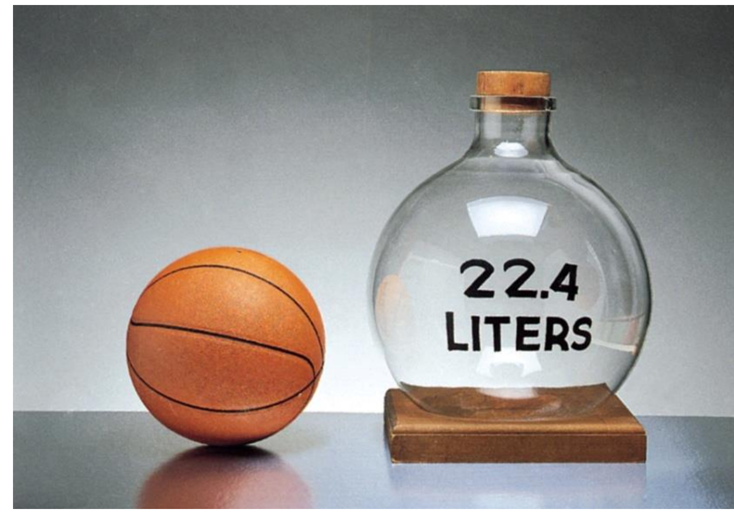
$1 \text{ mole } NH_3 \longrightarrow 1 \text{ mole } NO$

At constant *T* and *P*

$1 \text{ volume } NH_3 \longrightarrow 1 \text{ volume } NO$

The conditions 0 ^{0}C and 1 atm are called **standard temperature and pressure (STP).**

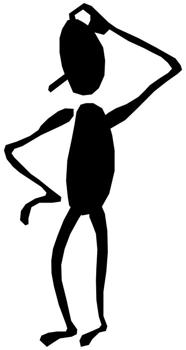
Experiments show that at STP, 1 mole of an ideal gas occupies 22.414 L.



$$PV = nRT$$

$$R = \frac{PV}{nT} = \frac{(1\text{ atm})(22.414\text{L})}{(1\text{ mol})(273.15\text{ K})}$$

$$R = 0.082057\text{ L} \bullet \text{atm} / (\text{mol} \bullet \text{K})$$



What is the volume (in liters) occupied by 49.8 g of HCl at STP?

T = 0 ^{0}C = 273.15 K

P = 1 *atm*

$$n = 49.8 \text{ g} \times \frac{1 \text{ mol HCl}}{36.45 \text{ g HCl}} = 1.37 \text{ mol}$$

$PV = nRT$

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

$$V = \frac{1.37\ \cancel{\text{mol}} \times 0.0821 \frac{\text{L}\cdot\cancel{\text{atm}}}{\cancel{\text{mol}\cdot\text{K}}} \times 273.15\ \cancel{\text{K}}}{1\ \cancel{\text{atm}}}$$

V = 30.6 L

Density (d) Calculations

$$d = \frac{m}{V} = \frac{P\mathcal{M}}{RT}$$

m is the mass of the gas in g

$\mathcal{M}$ is the molar mass of the gas

Molar Mass ($\mathcal{M}$) of a Gaseous Substance

$$\mathcal{M} = \frac{dRT}{P}$$

d is the density of the gas in g/L

Ideal Gas Equation

Boyle's law: $V \propto \frac{1}{P}$ (at constant n and T)

Charles' law: $V \propto T$ (at constant n and P)

Avogadro's law: $V \propto n$ (at constant P and T)

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

$$V = \text{constant} \times \frac{nT}{P} = R\frac{nT}{P}$$

R is the **gas constant**

$$\boxed{PV = nRT}$$

Determine Pressure:

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

Determine the pressure in atm exerted by 12.0 grams of sulfur dioxide gas at 22. ^{0} C in a 750.0 ml vessel.

1. Change all units:

$$n\ SO_2 = 12.0 \text{ grams of } SO_2 \left\| \frac{1 \text{ mole } SO_2}{64.0 \text{ g}} \right\| = 0.188 \text{ moles } SO_2$$

T= 22 + 273 = 295 K

V= 0.750 L

2. Plug into equation

$$P = \frac{(0.188 \text{ moles})(0.0821)(295 \text{ K})}{0.750 \text{ L}} =$$

6.07 atm

Determine Volume:

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

Determine the volume occupied by 5.00 grams of sulfur dioxide at 10.^{0}C exerting a pressure of 12.5 Kpa.

1. Change all units:

$$n\ SO_2 = 5.00 \text{ grams of } SO_2 \left| \frac{1 \text{ mole } SO_2}{64.0 \text{ g}} \right| = 0.0781 \text{ moles } SO_2$$

T= 10. + 273 = 285 K

$$P = 12.5 \left| \frac{1 \text{ atm}}{101.3\text{KPa}} \right| = 0.123 \text{ atm}$$

2. Plug into equation

$$V = \frac{(0.0781 \text{ moles})(0.0821)(285 \text{ K})}{0.123 \text{ atm}} =$$

14.8 L

Determine Temperature:

$$T = \frac{PV}{n\,R}$$

Determine the temperature in Celsius of a 2.50 liter vessel that contains 14.00 grams of sulfur dioxide and exerts a pressure of 255.2 Kpa.

1. Change all units:

$$n\ SO_2 = 14.00 \text{ grams of } SO_2 \left| \frac{1 \text{ mole } SO_2}{64.0 \text{ g}} \right| = 0.219 \text{ moles } SO_2$$

P = 255.2 / 101.3 KpPa= 2.52 atm

2. Plug into equation

$$T = \frac{(2.52\text{ atm})(2.50\text{ L})}{(0.0821)(0.219\text{ moles})} =$$

350. K

77. celsius

Determine sample size:

$$n = \frac{PV}{RT}$$

Determine the mass of a sample of sulfur dioxide in a 750.0 milliliter vessel that exerts a pressure of 875 mmHg at 10.^{0}C.

1. Change all units:

P $\quad 875 \text{ mmHg} = \frac{1 \text{ atm}}{760 \text{ mmHg}} = 1.15 \text{ atm}$

T= 10.+ 273 = 283 K

750.0 ml = 0.7500 L

2. Plug into equation

$$n = \frac{(1.15\,\text{atm})(0.7500\,\text{L})}{(0.0821)(283\,\text{K})} = \quad 0.0371 \text{ moles}$$

3. Solve for grams n X mm

$0.0371 \text{ moles } SO_2 \quad \frac{64.0 \text{ g } SO_2}{1 \text{ mole}} = \quad 2.38 \text{ g } SO_2$

Determine the Gas in a vessel:

$$n = \frac{PV}{RT}$$

A label has fallen off a **550.0 milliliter** vessel that exerts a pressure of **775 mmHg at 20.^{0}C.** The empty cylinder has a mass of **987.00 grams.** The cylinder with the gas in it has a mass of **988.49** grams.. Determine the MM of the gas.

1. Change all units:

P 775 mmHg $= \left| \frac{1 \text{ atm}}{760 \text{ mmHg}} \right| =$ 1.02 atm

T= 20.+ 273 = 293 K

550.0 ml = 0.5500 L

2. Plug into equation

$$n = \frac{(1.02\ \text{atm})(0.550\ \text{L})}{(0.0821)(293\ \text{K})} =$$

0.0233 moles

3. Solve for MM by dividing grams of sample by moles

$$\frac{1.49\ \text{g}}{0.0233\ \text{mole}} =$$

63.9 g/mole

Could the gas be sulfur dioxide?

Gas Stoichiometry

Gases are affected by temperature and pressure and must be handled differently.

Gases are affected the same way therefore in a chemical reaction in which the reactants and products are gases the volumes the molar volume of a gas may be used as a conversion factor

Gas Stoichiometry

- Many chemical reactions involve gases as a reactant or a product
- **Gas Stoichiometry –** the procedure for calculating the volume of gases as products or reactants
- Gases also have a molar volume (L/mol) rather than concentration.
- This is the conversion factor used to convert (litres of gas) to (moles of gas)
- The Ideal Gas Law ($PV = nRT$) may also be required to:

    A) find the number of moles of reactant

    B) Find the V, P, or T of the product

STP = 22.4L/mol

Molar Volume

- Molar volume is the **same** for all gases at the same temperature and pressure
- Remember, all gases have the same physical properties!
  - At STP, molar volume = 22.4 L/mol (101.325 kPa and 0°C)

This can be used as a conversion factor just like molar mass!

At STP, one mole of gas has a volume of 22.4 L, which is approximately the volume of 11 “empty” 2 L pop bottles.

Example # 1

- What volume of ammonia gas can be produced at 22.0°C and 787 mmHg given 25.0 milliliters of each reagent?

- First write the balanced equation!

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

- Note: the coefficients represent equivalent volumes of the gases at the same temperature and pressure.

- Next determine the limiting reagent:

$$25.0\text{ mL } N_2 \left| \frac{2\ NH_3}{1\ N2} \right| = 50.0\text{ mls } NH_3\,(g)$$

$$25.0\text{ ml } H_2 \left| \frac{2\ NH_3}{3\ H2} \right| = 16.7\text{ mls } NH_3\,(g)$$

How much leftover ???

Example #2

- Hydrogen gas is produced when sodium metal is added to water. What mass of sodium is necessary to produce 20.0L of hydrogen at STP?
- Remember: 22.4L/mol for STP

$$2Na_{(s)} \quad + \quad 2H_2O_{(l)} \rightarrow 2NaOH_{(aq)} + H_{2(g)}$$

m = ? — V = 20.0L

22.99g/mol — 22.4L/mol

$$20.0\cancel{L} \times \frac{1\ m\cancel{o}l}{22.4\ \cancel{L}} \times \frac{2mol\ Na}{1\ mol\ H} \times \frac{22.99g}{1\ \cancel{m}ol\ Na} = 41.1\ g\ Na_{(s)}$$

****Remember – molar volume is the conversion factor for gases just like molar mass is the conversion factor in gravimetric stoichiometry**

Example #3

- If 300.0g of propane burns in a gas barbecue, what volume of oxygen measured at STP is required for the reaction?
- Remember: 1 mol of any gas occupies 22.4 L at STP

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

m = 300g   m = ?

44.11g/mol   22.4L/mol

$$300.0\ \cancel{g} \times \frac{1\ \cancel{mol}}{44.11\ \cancel{g}} \times \frac{5\ mol\ O_2}{1\ \cancel{mol}\ C_3H_8} \times \frac{22.4\ L}{1\ mol} = 762\ L\ O_{2(g)}$$

****Remember – molar volume is the conversion factor for gases just like molar mass is the conversion factor in gravimetric stoichiometry**

Gas Stoichiometry Summary

1. Write a balanced chemical equation and list the measurements, unknown quantity symbol, and conversion factors for the measured and required substances.

2. Convert the measured quantity to a chemical amount using the appropriate conversion factor

3. Calculate the chemical amount of the required substance using the mole ratio from the balanced chemical equation.

4. Convert the calculated chemical amount to the final quantity requested using the appropriate conversion factor.

Given Volume Going To Grams Non- STP

1. Solve PV= nRT for moles A

   n_A= PV/ RT

2. Solve Stoich Expression for Grams B

Moles A	# Moles B	MM B
	# Moles A	1 Mole B

Example

If the conditions are <u>not</u> STP , the molar volume <u>cannot</u> be used! You must use the ideal gas law to find the gas values using moles determined from stoichiometry

- 5.50×10^3 liters of ammonia formed at 450kPa and 80.°C. How many grams of hydrogen were required?

$$2N_{2(g)} + 3H_{2(g)} \rightarrow 2NH_{3(g)}$$

m = X Grams	V= 5.50 x 103 l
M = 2.02 g/mol	P = 450kPA
T = 353.13K	

$$n = \underline{PV/RT} = \frac{(4.44\text{atm})(5.50 \times 10^3 \text{ liters})}{(0.0821^{\text{atm}\bullet\text{L}}/_{\text{mol}\bullet\text{K}})(4.44\text{atm})}$$

$$= \rightarrow 6.70 \times 10^4 \text{ moles } \mathbf{NH_{3(g)}}$$

$$6.70 \times 10^4 \text{ moles } \mathbf{NH_{3(g)}} \left| \frac{3\ H_2}{2\ NH_3} \right| \frac{2.02 \text{ grams}}{1 \text{ mole } H_2} \right| = 203 \text{ kilograms} H_{2(g)}$$

Example

If the conditions are not STP , the molar volume cannot be used! You must use the ideal gas law to find the gas values using moles determined from stoichiometry

- What volume of ammonia at 450kPa and 80°C can be obtained from the complete reaction of 7.5kg of hydrogen with nitrogen?

$$2N_{2(g)} + 3H_{2(g)} \rightarrow 2NH_{3(g)}$$

H_2: m = 7500g, M = 2.02 g/mol

NH_3: m = ?, P = 450kPA, T = 353.13K

$$7500\,\cancel{g} \times \frac{1\ \cancel{mol}}{2.02\ \cancel{g}} \times \frac{2\ NH_3}{3\ H_2} = 2475.2475\ \text{mol}\ NH_{3(g)}$$

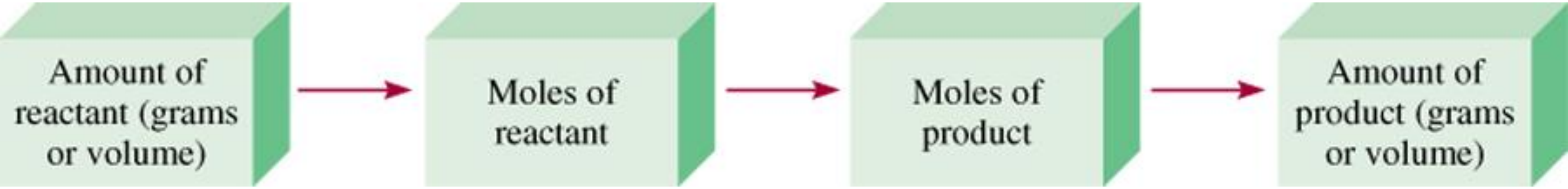
$$PV = nRT \rightarrow V = \frac{nRT}{P} = \frac{(2475.2475\ \cancel{mol})(0.0821^{atm\bullet \cancel{L}}/_{\cancel{mol}\bullet K})(353.1\cancel{5}K)}{(4.44\cancel{atm})}$$

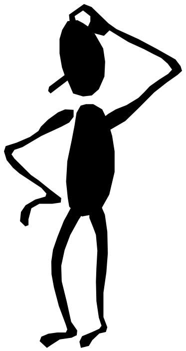
$$= 16150.10L \rightarrow \mathbf{1.6 \times 10^4\ L\ of\ NH_{3(g)}}$$

Given Grams going to Volume of a gas non- STP

- 1. Solve stoich expression to change grams of material A to moles of gas B
- $$\text{Grams A} \left| \frac{1 \text{ Mol A}}{\text{MM A}} \right| \frac{\# \text{ moles B}}{\# \text{moles A}} \right| = \text{Moles B}$$
- 2. Solve PV=nRT for V_B $$V_B = \frac{n_B RT}{P}$$

Gas Stoichiometry





What is the volume of CO_2 produced at 37^0 C and 1.00 atm when 5.60 g of glucose are used up in the reaction:

$$C_6H_{12}O_6\ (s) + 6O_2\ (g) \longrightarrow 6CO_2\ (g) + 6H_2O\ (l)$$

$$\text{g } C_6H_{12}O_6 \longrightarrow \text{mol } C_6H_{12}O_6 \longrightarrow \text{mol } CO_2 \longrightarrow V\ CO_2$$

$$5.60\ \cancel{\text{g } C_6H_{12}O_6} \times \frac{1\ \cancel{\text{mol } C_6H_{12}O_6}}{180\ \cancel{\text{g } C_6H_{12}O_6}} \times \frac{6 \text{ mol } CO_2}{1\ \cancel{\text{mol } C_6H_{12}O_6}} = 0.187 \text{ mol } CO_2$$

$$V = \frac{nRT}{P} = \frac{0.187\ \cancel{\text{mol}} \times 0.0821\ \dfrac{\text{L}\cdot\cancel{\text{atm}}}{\cancel{\text{mol}}\cdot\cancel{\text{K}}} \times 310.15\ \cancel{\text{K}}}{1.00\ \cancel{\text{atm}}} = 4.76 \text{ L}$$

Gas Stoichiometry #1

If reactants and products are at the same conditions of temperature and pressure, then mole ratios of *gases* are also volume ratios.

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

3 moles H_2 + 1 mole N_2 → 2 moles NH_3

3 liters H_2 + 1 liter N_2 → 2 liters NH_3

Gas Stoichiometry #2

How many liters of ammonia can be produced when 12 liters of hydrogen react with an excess of nitrogen?

$$\underline{3}\ H_2(g) + N_2(g) \rightarrow \underline{2}NH_3(g)$$

$$\frac{12\ \cancel{L\ H_2}}{} \left| \frac{2\ L\ NH_3}{3\ \cancel{L\ H_2}} \right. = 8.0\ L\ NH_3$$

Gas Stoichiometry #3 At STP

How many liters of oxygen gas, at STP, can be collected from the complete decomposition of 50.0 grams of potassium chlorate?

$$2\ KClO_3(s) \rightarrow 2\ KCl(s) + 3\ O_2(g)$$

$$\frac{50.0\ \cancel{g\ KClO_3}}{} \times \frac{1\ \cancel{mol\ KClO_3}}{122.55\ \cancel{g\ KClO_3}} \times \frac{3\ \cancel{mol\ O_2}}{2\ \cancel{mol\ KClO_3}} \times \frac{22.4\ L\ O_2}{1\ \cancel{mol\ O_2}}$$

$$= 13.7\ L\ O_2$$

Gas Stoichiometry #4

How many liters of oxygen gas, at 37.0°C and 0.930 atmospheres, can be collected from the complete decomposition of 50.0 grams of potassium chlorate?

$$2\ KClO_3(s) \rightarrow 2\ KCl(s) + 3\ O_2(g)$$

$$\frac{50.0\ \cancel{g\ KClO_3}}{} \cdot \frac{1\ \cancel{mol\ KClO_3}}{122.55\ \cancel{g\ KClO_3}} \cdot \frac{3\ mol\ O_2}{2\ \cancel{mol\ KClO_3}} = 0.612\ mol\ O_2$$

$$V = \frac{nRT}{P} = \frac{(0.612\ mol)(0.0821\ \frac{L \cdot atm}{mol \cdot K})(310\ K)}{0.930\ atm} = 16.7\ L$$

Gas Stoichiometry #5

15.5 liters of oxygen gas was collected at 37.0°C and 0.930 atmosphere when potassium chlorate decomposed. How many grams were used?

$$2\ KClO_3(s) \rightarrow 2\ KCl(s) + 3\ O_2(g)$$

$$n = \frac{PV}{RT} = \frac{(15.0\ L)\ 0.930\ atm}{(0.0821\ \frac{L \cdot atm}{mol \cdot K})\ (310K)} = 0.548\ mol\ O_2$$

$$0.548\ moles\ O_2 \left| \frac{2\ mol\ KClO_3}{3\ mol\ O_2} \right| \frac{122.55\ g\ KClO_3}{1\ mol\ KClO_3} = 44.8\ grams$$