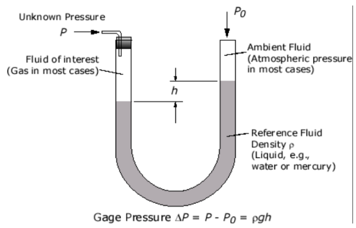
Chapter 5: Gases

**Pressure:**

Measured with a barometer or manometer



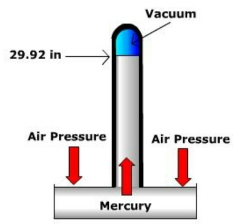

Units of Pressure:

1 atm = 760 torr = 760 mm Hg

= 101,325 Pa (101.325 kPa)

= 29.92 inches Hg

= 14.7 lb/in$^2$

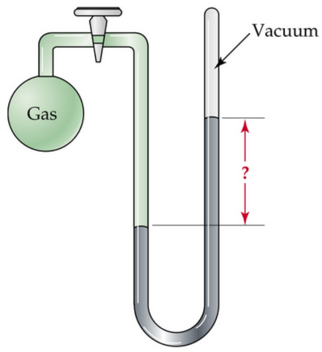



Ex: Convert 30.06 inches of Hg to atm

**Math with a sealed tube manometer:**

- used to measure pressures of gases that are below atm pressure

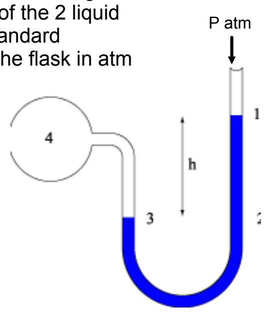
Ex: A sealed tube manometer has a gas in it, in which the (?) h value = 6.5 cm. Calculate the pressure in the flask (of the gas) in torr, pascals, and atm? The liquid in the tube is Hg.

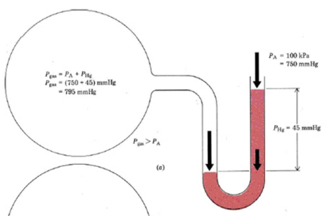



**Math with a open ended monometer:**

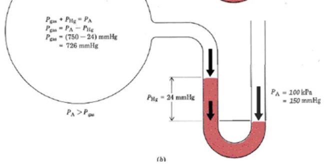
- used to measure measure larger pressure values of gases, in relation to atm pressure

Ex: An open-ended manometer is used to measure pressure of a gas. A gas is inserted into the flask and the pressure difference of the 2 liquid levels is measured to be 118 mm inches. If you are at standard atmopheric pressure, what is the pressure of the gas in the flask in atm and torr?



$P_{gas} > P_{atm}$ $P_{gas} = P_{atm} + h$



$P_{gas} < P_{atm}$ $P_{gas} = P_{atm} - h$

**Gas Laws:**

**Boyles:** As pressure __________ volume decreases

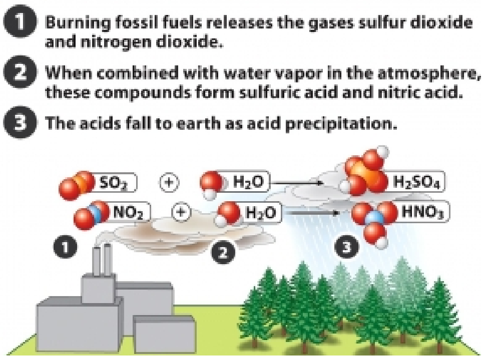
$P_1V_1 = P_2V_2$ (constant) $k = P*V$

http://www.youtube.com/watch?v=N5xft2fIqQU

a.) Only works perfectly at low pressures or for an ideal gas

b.) Deviations from projected values are small, so we still use it

Ex: Acid rain is formed when $SO_2$ reacts with $H_20$ in the atmosphere. If $SO_2$ has a volume of 1.53 L at $5.6 \times 10^3$ Pa, what will the volume of $SO_2$ be when Pressure increases to $1.5 \times 10^4$ Pa?




Ex: You have a bag of potato chips that you are taking in a hot air balloon ride. Before take-off the atm pressure is 29.4 inches Hg. You go to go up, and the pressure reduces to 26.8 inches Hg. Convert both units to atm pressure. If the bag had an initial volume of 2.30 L, and has maximum value of 3.00 L, will the chip bag explode as you go up in the hot air balloon?

**Charle's Law:**

- Volume and temperature are directly related $V = bT$ (b = constant)

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

V = volume

T = Kelvins

(Kelvins = C + 273)

http://www.youtube.com/watch?v=bWhs5L_gBTI

- Absolute zero -

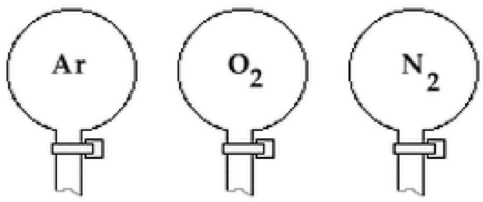
Ex: Sample of gas at 15 C and 1 atm = 2.58 L. If temp is changed to 38 C (and same pressure), what is the new volume.

**Avogadro's Law:**

- Equal volumes of gases at same temp and pressure contain the same number of particles

$V = a*n$ or $\frac{V_1}{n_1} = \frac{V_2}{n_2}$ V = volume n = moles

a = constant




| | | | |
|---|---|---|---|
| Volume: | 22.4 L | 22.4 L | 22.4 L |
| Mass: | 40 g | 32 g | 28 g |
| Quantity: | 1 mol | 1 mol | 1 mol |
| Pressure: | 1 atm | 1 atm | 1 atm |
| Temperature: | 273 K | 273 K | 273 K |

Ex:

You have a 12.2 L sample of oxygen with 0.50 mol at 1 atm and 25 C. If all oxygen ($O_2$) becomes ozone ($O_3$), what will the volume be?

**Ideal Gas Law:**

$PV = nRT$ $\qquad D = \dfrac{P*MM}{RT}$ $\qquad PV = \dfrac{gRT}{MM}$

P = atm R = 0.08206 (L*atm)/(K*mol) T = kelvins

V = Liters D = g/L n = moles MM = molar mass

***Note: You can actually use the ideal gas law for any gas law!!***

***HOW: Place the variables that "Change" on 1 side, and the "constant" variables on the other***

Let's try manipulation of the ideal gas law for the next problems:

**Ex:** An ammonia gas sample is at 7.0 mL and at P = 1.68 atm. The gas is compressed to 2.7 mL. Constant temperature is maintained. Calculate the new P?

**Ex:** The pressure of methane gas has a volume of 3.8 L at 5 C. It is heated to 86 C, find the new volume.

**Ex:** Dibromane gas ($B_2H_6$) has an initial P = 345 torr at a T = -15 C and a volume = 3.48 L. If T increases to 36 C and P increases to 468 torr, what is volume?

Ex: A sample of gas contains 0.35 mol Ar gas at T = 13 C and P = 568 torr. The sample is heated to 56 C and P increases to 897 torr. Calculate the change in Volume.

**Ideal Gas Law Practice:**

Ex: A sample containing 0.214 moles of a gas at 12.0 C occupies a volume of 1545. mL. What pressure does the gas exert?

Ex: You have a sample of oxygen gas at 37 C. If you inject .450 moles of it into a closed container with a volume of 5.0 L, what pressure would be exhibited?

Ex: The density of a gas was measured at 1.50 atm and 27 C. D = 1.95 g/L. Calculate its MM?

Ex: A 502.8-g sample of $X_{2(g)}$ has a volume 9.0 L at 10 atm and 102 °C. What is element X?

**Gas stoichiometry:**

Ways to use moles in gas stoich:

a.) IF AT 273 K and 1.00 atm (STP) 1 mol (any gas) = 22.4 L

b.) Use ideal gas law to calculate moles of gas

*** Either way, use stoichiometry with the mole value found or needed

Ex: Nitrogen has has a volume of 1.75 L at STP. What amount of moles are present?

Ex: Quicklime (CaO) is produced using thermal decomposition by

$CaCO_{3(s)}$ ----> $CaO_{(s)}$ + $CO_{2(g)}$

Calculate the volume of carbon dioxide at STP conditions from 152 g of $CaCO_3$ decomposing.

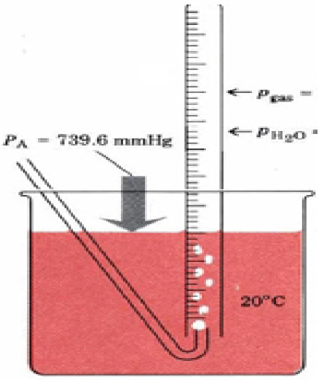
Ex: A sample of methane has a volume of 2.80 L at 25 C and 1.65 atm. It is mixed with oxygen gas that has a volume of 35.0 L and 31 C, at a pressure of 1.25 atm. The mixture was ignited and $CO_2$ and $H_2O$ were formed. Calculate the theoretical volume of $CO_2$ formed from this reaction at a P = 2.50 atm and a T = 125 C.

## Dalton's Law of Partial Pressure:

- In a mixture of gases, each gases pressure is independent of the rest, and they can all be added to = the total pressure

$P_T = P_1 + P_2 + P_3 ......$

$P_T = P_{gas} + P_{H20}$ Always find $P_{H20}$ by using table




Ex: Mixtures of He and $O_2$ are added to a scuba diving tank.

He = 46 L at 25 C and 1 atm and O2 = 12 L at 25 C and 1 atm. (Both gases in these quantity, temp and pressure were added)

Calculate the partial pressure of each gas and the $P_T$ in the tank at 25 C

**Mole fractions and gases:**

-Ratio of the number of moles of a given component in a mixture of gases to the total # of moles in the mixture.

$$\text{mole fraction} = X_1 = \frac{n_1}{n_{total}}$$

IF $\dfrac{n_1}{n_{total}} = \dfrac{P_1}{P_{total}}$ Then: $X_1 = \dfrac{P_1}{P_{total}}$ or $P_1 = X_1 * P_{total}$

Ex: Partial $P_{O2}$ = 156 torr in air with a total $P_T$ = 746 torr. Calculate the mole fraction of $O_2$

Ex: Mole fraction of $N_2$ in air is $X_{n2}$ = 0.7808. Calculate the partial pressure of $N_2$ in air when $P_{atm}$ = 760. torr

Ex: A sample of solid $KClO_3$ is heated to produce KCl and $O_2$. The $O_2$ produced was collected by water displacement and the water temp = 22 C, The $P_T$ = 754 torr. The volume of $O_2$ collected was 0.650 L. If vapor pressure of water at 22 C = 21 torr, calculate the partial pressure of $O_2$ and the mass of $KClO_3$ that was decomposed.

**Kinetic Molecular Theory and Gases:**

**Four postulates of the KMT theory: (ideal gases)**

1.) Particles are so small, V = 0 for an individual particle

2.) Particles are in constant motion, their collisions with walls of containers = pressure

3.) Particles exert no forces upon each other (neither attraction or repulsion)

4.) Average KE of gas particles is directly proportional to the temp (kelvins)

BUT REAL GASES DO NOT FIT THIS THEORY!!!

WHY??

**Real gases.... behaving like ideal?**

- No real gases behave EXACTLY like ideal
- ***Low pressure*** and ***high temp*** will cause the closest behavior of real to ideal

Why????

Relative attractive forces that gases have toward each other:

$H_2$ < $N_2$ < $CH_4$ < $CO_2$

This means.... Hydrogen is most like ideal gas

**Kinetic energy and Temperature of gases:**

Kinetic energy increases as temperature increases

Predict exact KE as follows:

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

Ex: Consider three samples of gas, 1.0 L each, and all at STP.

$H_2$ Xe $Cl_2$ $O_2$

a.) Rank the gases in order of increasing average KE

b.) Rank the gases in order of increasing velocity

**Root Mean Square Velocity:**

$$U_{rms} = \sqrt{\frac{3RT}{M}}$$

R = 8.3145 J/(K*mol) $U_{rms}$ = m/s

M = kg mass of a mole of the gas

**Mean Free Path-** Average distance a gas particle can travel before geing in a collision

Ex: Calculate the root mean square velocity for atoms of a sample of He gas at 25 C.

Ex: Calculate the RMS veleocity of atoms in a sample of oxygen gas at
a.) 0 C and b.) 300. C

**Diffusion and Effusion:**

**Diffusion -** Mixing of gases (like perfume being sprayed in by a classmate)

**Effusion** - Passage of gas molecules through a tiny hole into an evacuated container

**Grahams Law of Effusion:**

$$\frac{\text{Rate of effusion for gas 1}}{\text{Rate of effusion for gas 2}} = \frac{\sqrt{M_2}}{\sqrt{M_1}}$$

M = molar mass

*** Note: Place heavier gas on top, the answer you receive will tell you how much "faster" the smaller gas is moving

Ex: Calculate the rate of effusion of $H_2$ has and $UF_6$.

**Predicting rates of diffusion:**

- Smaller mass gases move faster (diffuse), but it is hard to predict an exact speed because of their unpredictable collisions with other gases in their path.

Only predictable rule: Lower MM gases = Increase diffusion speed

**Chemistry in the atmosphere:**

What gases make-up our atmosphere?

Layers of atmosphere:

a.) Ozone (very high level)

b.) Troposphere

Concerns of what chemistry is happening in each of these?

**Ozone** = protects us from UV radiation, it is being destroyed

**Troposphere** = Harboring pollution (gases and particles)

**Biggest concerns to troposphere:**

**Burned petroleum** - puts off CO, $CO_2$, NO and $NO_2$

#1 = NO: (Emitted from car engines)

NO ---> $NO_2$ (NO is oxidized immediately in air)

$NO_2$ ---> NO + O (That $NO_2$ then breaks up)

O + $O_2$ ---> $O_3$ (Ozone is made, and it is very reactive)

$O_3$ ---> $O_2$ + O*(*= really reactive)

O* + $H_2O$ ---> OH-

Then OH- + $NO_2$ ----> $HNO_3$

Also, Photochemical smog is made in troposphere:

$NO_2$ ---> NO + O

O + $O_2$ ---> $O_3$

NO + 1/2 $O_2$ ----> $NO_2$ +

3/2 $O_2$ ---> $O_3$

#2 Concern: SO2 from coal burning

$S_{(in\ coal)}$ + $O_2$ ---> $SO_{2(g)}$

Then: $2SO_{2(g)}$ + $O_2$--- > $2SO_{3(g)}$

$SO_{3(g)}$ + $H_2O_{(l)}$ ---> $H_2SO_{4\ (l)}$ (ACID RAIN!!!)

**Concerns of acid rain:**

1.) Acidifies fresh water lakes and streams

2.) As we continue to try to reduce petroleum use, coal use for electricity is going up!

3.) Low S coal, high S coal exists. BUT we are running out of Low S coal!!

4.) Use of scrubbers to clean the exhaust air?

How scrubbers work:

$CaCO_{3(s)}$ ---> $CaO_{(s)}$ + $CO_{2(g)}$

$CaO_{(s)}$ + $SO_{2(g)}$ ---> $CaSO_{3(s)}$

The $CaSO_{3(s)}$ is then cleaned out with aqueous lime solution and a thick slurry solution is washed out.

But the $CaSO_{3(s)}$ slurry is a problem.... we have nothing to do with it!!