Chemistry 6A F2007

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Boyle's Law: *Pressure vs. Volume*

as pressure increases, volume decreases!

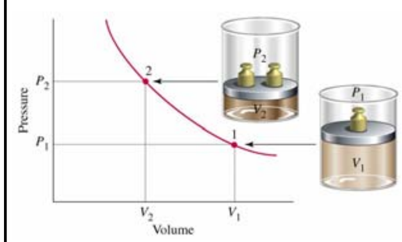
$$V \propto \frac{1}{P}$$

or

$$V \times P = \text{Constant}$$

$$V_1 \times P_1 = V_2 \times P_2$$

$$V_2 = \frac{V_1 \times P_1}{P_2}$$



if $P_2 = 2 \times P_1$

$$V_2 = \frac{V_1 \times P_1}{2P_1} = \frac{1}{2} V_1$$

double the pressure, halve the volume

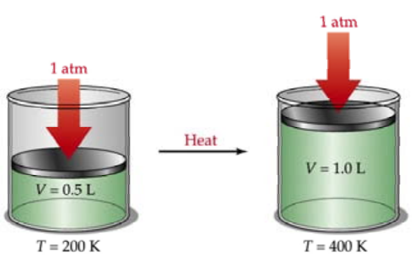
Charles's Law: ***Volume vs. Temperature (absolute)***

$$V \propto T$$

or

$$\frac{V}{T} = \text{Constant}$$

as temperature increases, volume increases!



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

if $T_2 = 2 \times T_1$

$$V_2 = \frac{V_1}{T_1} \times 2T_1 = 2 \times V_1$$

Double the ***Kelvin*** temperature, double the volume!

Calculate the new volume of a 352 mL sample of helium at a 25.0 ^{o}C after the temperature is changed to 50.0 ^{o}C.

$$\frac{V_{(1)}}{T_{(1)}} = \frac{V_{(2)}}{T_{(2)}} \qquad V_{(2)} = \frac{V_{(1)} \times T_{(2)}}{T_{(1)}}$$

The temperature doubles so the volume doubles, ***Right?***

Only if one uses the Kelvin scale!

$$V_{(2)} = \frac{352\ \text{mL} \times (50.0 + 273.15)\text{K}}{(25.0 + 273.15)\text{K}} = 382\ \text{mL}$$

What temperature (in °C) would change the volume of 35.2 mL of helium at 25.0 °C to 575 μL?

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

$$T_2 = \frac{575\ \mu L \times (25.0 + 273.15)K}{35.2\ mL \times \frac{L}{10^3\ mL} \times \frac{10^6\ \mu L}{1 L}} - 273.15$$

$$= -268\ ^oC$$

General Gas Law:

Combining Charles's and Boyle's Laws…

$$\frac{V}{T} = \text{Constant} \qquad V \times P = \text{Constant}$$

So at two sets of conditions:

$$\frac{P_1 \times V_1}{T_1} = \text{a constant} = \frac{P_2 \times V_2}{T_2}$$

$$\boxed{\frac{P_1 \times V_1}{T_1} = \frac{P_2 \times V_2}{T_2}}$$

Example: A gas sample has a volume of **2.50 L** when it is at a temperature of **30.0°C** and a pressure of **1.80 atm**.

What volume in liters will the sample have if the pressure is increased to **3.00 atm**, and the temperature is increased to **100.0°C**?

Solution: The problem can be solved using the combined gas law.

- First, identify the initial and final conditions.
- Be sure all like quantities are in the same units.
- Express the temperatures in Kelvin

	initial	final
P		
V		? L
T	+ 273.15	+ 273.15

Example: A gas sample has a volume of 2.50 liters when it is at a temperature of 30.0°C and a pressure of 1.80 atm.

What volume in liters will the sample have if the pressure is increased to 3.00 atm, and the temperature is increased to 100°C?

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

Substituting in the values given:

T must be in Kelvin!

$$V_2 = \frac{(100.0 + 273.15)K \times 1.80\ \text{atm} \times 2.50\ L}{3.00\ \text{atm} \times (30.0 + 273.15)K} = 1.85\ L$$

Example: What is the volume (in ml) of a gas initially at 30.0 °C, 775 torr and 3.25 L that is changed to 0.555 atm and 25.0 °C?

Solution:

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

Solving for the new volume: $$V_2 = \frac{P_1 \times V_1}{T_1} \times \frac{T_2}{P_2}$$

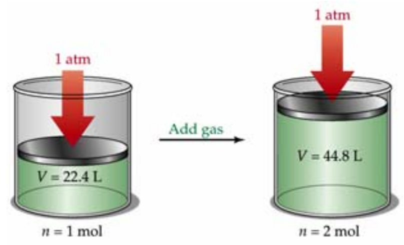
Substituting in the values given:

$$V_2 = \frac{775 \text{ torr} \times 3.25 \text{ L} \times \frac{10^3 \text{ ml}}{1 \text{ L}} \times (25.0 + 273.15)\text{K}}{0.555 \text{ atm} \times \frac{760 \text{ torr}}{\text{atm}} \times (30.0 + 273.15)\text{K}} = \underline{5870 \text{ ml}}$$

Avagadro's Law: *Volume vs. number of particles (moles)*

Avagadro's Law:

Equal amounts of gases (moles) at the same temperature (T) and pressure (P) occupy equal volumes (V).

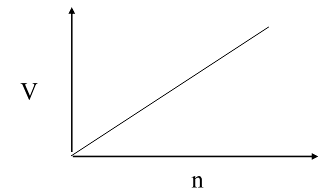


or…

$V \propto n$ or $\frac{V}{n} = \text{Constant}$

A plot of V vs. n yields a straight line

$y = mx + b$



Combining Avagodro's Law with the general gas law…

This in known as the "***Ideal Gas Law***"

$$\frac{P \times V}{n \times T} = \text{constant}$$

$$P \times V = n \times R \times T$$

$$R = \text{"gas constant"} = 0.08206 \frac{L \cdot atm}{mol \cdot K}$$

$$PV = nRT$$

Rules for Ideal Gas Law Calculations:

- Always convert the temperature to Kelvin (K = 273.15 + ^oC)
- Convert from grams to moles if necessary.
- Be sure to convert to the units of "R" (L, atm, mol & K).

Types of Ideal Gas Law problems you may encounter:

•Determination one unknown quantity of one gas variable (P, V, T, or n) given the other values directly or indirectly.

•Determine the new values of P, V, T, or n after a gas undergoes a change.

•Stoichiometry problems.

•Gas density and molar mass problems.

What is the volume (L) of 1.00 mole of any gas at STP?

STP = 1 atm and 0 °C (273.15 K)

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

Since the units of R: $\frac{L \cdot atm}{mol \cdot K}$

$$V = \frac{1.00\ \cancel{mol} \times 0.08206\ \frac{L \cdot \cancel{atm}}{\cancel{mol} \cdot \cancel{K}} \times 273.15\ \cancel{K}}{1.00\ \cancel{atm}} = 22.4\ L$$

Standard Molar Volume!

How many grams of hydrogen and nitrogen are needed to produce 15.0L of ammonia at ***STP***?

$$3H_2(g) + N_2(g) \longrightarrow 2NH_3(g)$$

$$15.0\ \cancel{L}\ NH_3 \times \frac{3\ \cancel{L}\ H_2}{2\ \cancel{L}\ NH_3} \times \frac{1\ \cancel{mol}\ H_2}{22.4\ \cancel{L}\ H_2} \times \frac{2.02\ g\ H_2}{1\ \cancel{mol}\ H_2}$$

$$= 2.03\ g\ H_2$$

Similarly, it takes 9.38g of N_2 to produce the required 15.0L of ammonia.

3.62 L of a gas at STP as a mass of 7.11g.
What is the molecular weight of the gas?

Recall that the ***molar mass*** (aka. ***molecular weight***) is the ratio of mass to moles

Using the ***STP*** volume on can find moles, so one can write:

$$M\ wt. = \frac{7.11\ g}{3.62\ \cancel{L} \times \frac{1\ mol}{22.4\ \cancel{L}\ H_2}} = 44.0\ g/mol$$

Lecture Problem: The density of a gas at STP is 1.18g/L. Calculate the molecular weight of the gas.

Recall that the ***molar mass*** (aka. ***molecular weight***) is the ratio of mass to moles

Using the ***STP*** volume on can find moles, so one can write:

$$\frac{1.18\ g}{1\ \cancel{L}} \times \frac{22.4\ \cancel{L}}{1\ mol} = 26.4\ g/mol$$