Gases

Background

Discussion: What do we already know about gases? Many of the concepts to be covered are based on the interpretation of 'everyday' experiences.

Macroscopic View: Microscopic View:

Lower pressure   Higher pressure

1 atm = 760 mmHg (Torr) = 101 kPa

  The behavior of a gas can be explained in terms of:

1. **Its VOLUME (in liters)**
2. **Its PRESSURE (most often in atm.)**
3. **Its TEMPERATURE (in Kelvin)**
4. **The number of MOLES (n) of gas present**

 Understanding gases is all about understanding the relationships between the VOLUME, PRESSURE, TEMPERATURE and # MOLES (n) of a gas sample.
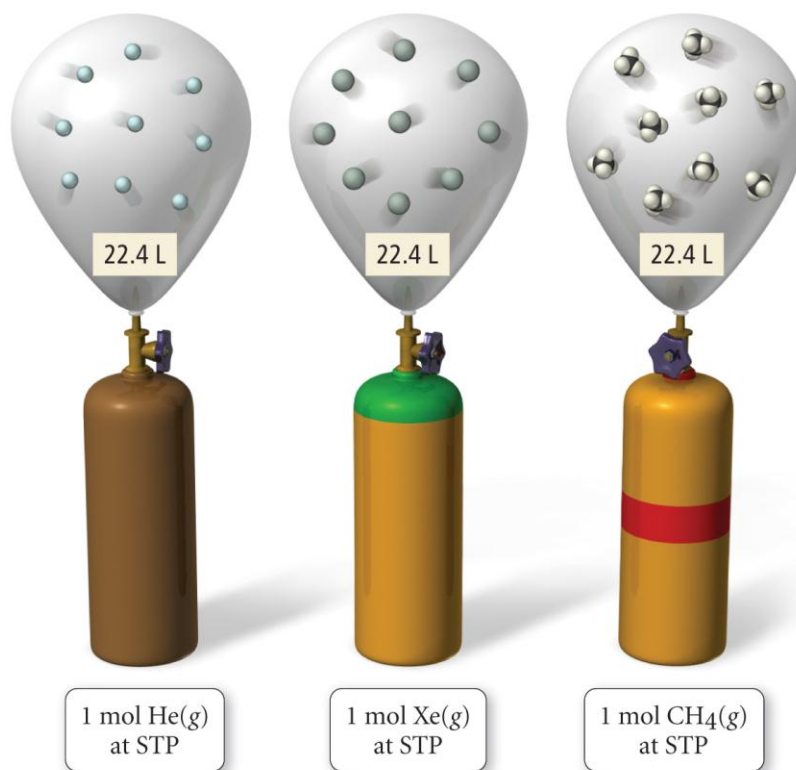
The Gas Laws

Avogadro's Law: Relationship between **# MOLES** of gas and **VOLUME** (P and T fixed)

Everyday observation: What takes up more space, 1 mole of gas or 2 moles of gas (assume constant temp. and pressure)?

$V \propto n$ (# moles)

At standard temperature (273 K) and pressure (1.0 atm.), 1 mole of ANY gas occupies 22.4 L.

1 mole = 22.4 L at STP

| 22.4 L | 22.4 L | 22.4 L |

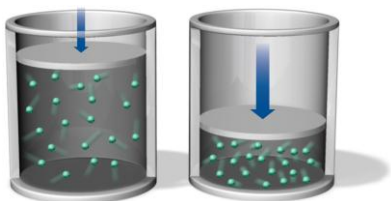| 1 mol He(g) at STP | 1 mol Xe(g) at STP | 1 mol CH₄(g) at STP |

<u>Examples</u>:

1. What volume does 3.5 moles of H_2 (g) occupy at STP?

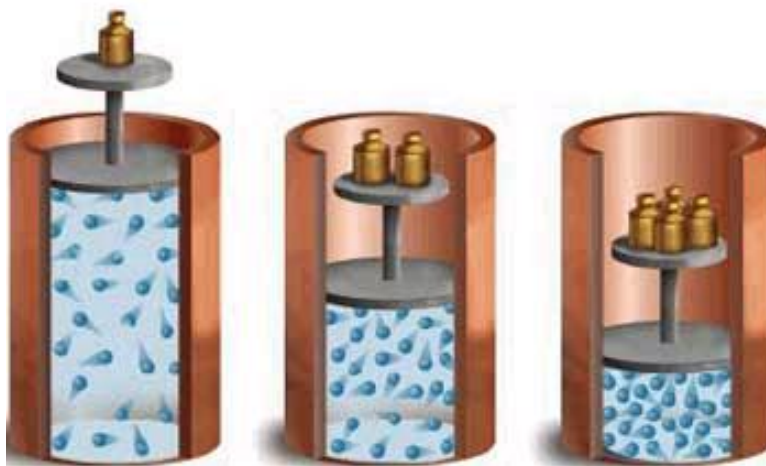2. How many atoms of He are there in a 5.0 L party balloon at STP?

Boyles Law:	Relationship between **PRESSURE** and **VOLUME** (n and T fixed)

Everyday observation: If you compress ('squeeze') the container a gas is in, does the gas pressure increase or decrease (assume fixed T and n).
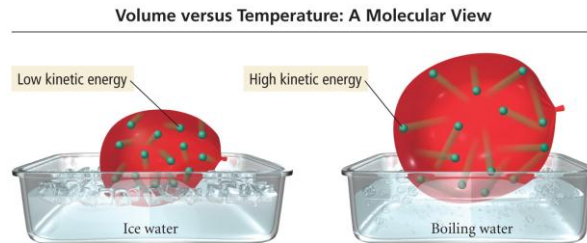
Boyle's Law

$$P \propto 1/V \quad or \quad PV = constant \text{ (table)}$$

Example: A sample of gas has a pressure of 2.0 atm. If the volume of the container the gas sample is in is decreased by a factor of 10, what is the new pressure of the gas inside? Assume T and n constant.
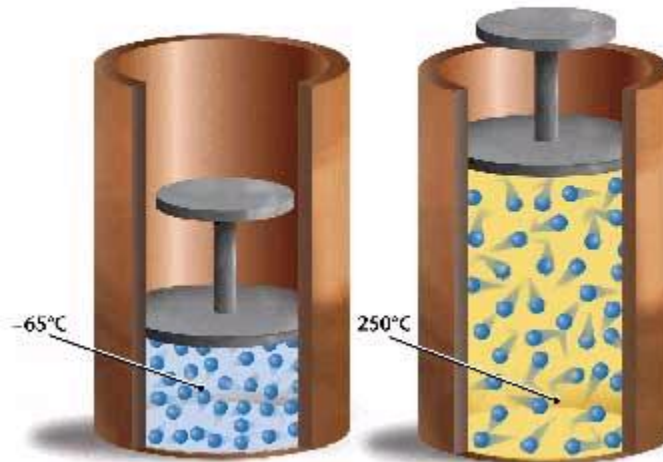
Charles' Law:	Relationship between:
	TEMPERATURE and **VOLUME** (n and P fixed) *or*
	TEMPERATURE and **PRESSURE** (n and V fixed)*

Everyday observation: Why does a hot air balloon rise? What time of day is best for ballooning?

Volume versus Temperature: A Molecular View

Low kinetic energy

High kinetic energy

Ice water

Boiling water

Charles's Law

−65°C

250°C

$$V \propto T$$

Example: A sample of gas occupies 12 L at 31°C. What volume does it occupy at 62°C? Assume P and n are constant.

The Ideal gas Law

> The three gas law equations can be combined to make a new equation (the **IDEAL GAS LAW**) that can be used to solve ANY 'static*' gas problem

Derivation of the ideal gas law (see appendix)

$$PV = nRT$$

Where: **R = 0.08206 L atm /mol K if the atm. pressure unit is used**
or   R = 8.314 /mol K if the Pa or N/m^2 pressure unit is used

Example: Calcium carbonate decomposes when heated to give solid calcium oxide and carbon dioxide gas. If 250 mL of CO_2 (g) is collected at 31°C and 1.3 atm pressure, then how many moles of CO_2 (g) is collected?

> Write down what you are given, *then* solve for the unknown quantity in 'static' problems

$P =$
$V =$
$n =$
$R =$
$T =$

Dynamic Problems: changing P, V or T for a fixed number of moles of gas

Derivation of the 'Dynamic' (before and after) gas law

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

(before)   (after)

Example: The pressure inside an aerosol can is 1.5 atm. at 25°C. What would the pressure inside the can be at 450°C? *Is there a moral to the story?*

As with the 'static' problems, write down what you are given *then* solve for the unknown quantity in 'dynamic' problems
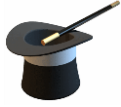
$P_1 =$ $P_2 =$

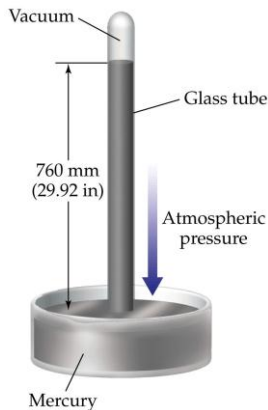$V_1 =$ $V_2 =$

$T_1 =$ $T_2 =$

(before)   (after)

Measuring Atmospheric and/or Gas Pressure

Statement: The mmHg pressure unit sounds 'all wrong' – isn't mm a length unit?

Reply: Yes, mmHg is a length unit, but it is related to the pressure of gas as measured by the height of a mercury *barometer* or *manometer* column (where 1 atm = 760mmHg)

Discussion: How does a mercury barometer work?

Discussion: Why is mercury used in barometers and manometers – why not water or some other 'safe' liquid?? Hint: what physical property is greatly different for Hg compared to other liquids?

**The height of a column of liquid used to measure pressure and it's density are related through the equation:**

$$P = gdh$$

Where P = pressure, g = gravitational constant $(9.81 ms^{-2})$ and d = density of the liquid in the column.

Interesting question: Suppose water was used in place of mercury in all manometers and barometers. What then would be standard atmospheric pressure (760mmHg) be if measured in mmH₂O? The density of water is 1.000 g/mL, the density of Hg = 13.596 g/mL).
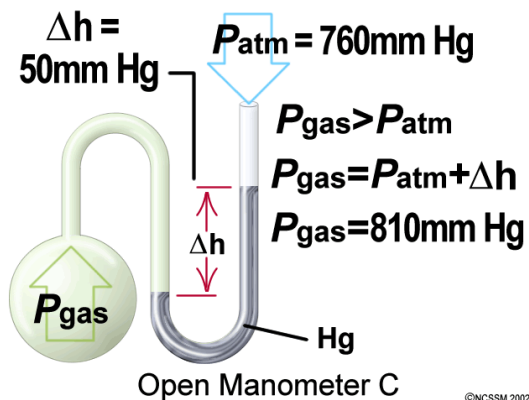
Ans: 10.33m

The mercury manometer

P_{atm}

$P_{gas}=P_{atm}=$ 760mm Hg

Hg

Open Manometer A

©NCSSM 2002
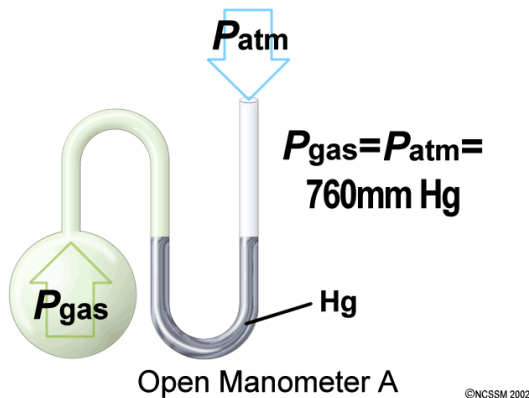
'Before' The pressure of the gas in the bulb and the atmospheric pressure are equal:

$$P_{gas} = P_{atm} = 760mmHg$$

Therefore, the Hg in the U-tube is at the same level on the 'gas side' and the 'atmosphere side'.

'Change': The pressure of the gas in the bulb is increased (how?). The extra pressure forces the Hg column down on the 'gas side' and 'up' on the 'atmosphere side' (like drinking a soda?!).

The *extra* pressure exerted by the gas is measured as a difference in height (Δh = 50 mm) between the respective Hg levels

Δh = 50mm Hg

P_{atm} = 760mm Hg

$P_{gas}>P_{atm}$
$P_{gas}=P_{atm}+\Delta h$
$P_{gas}=810mm\ Hg$

Δh

P_{gas}

Hg

Open Manometer C

©NCSSM 2002

After: The new pressure of the gas is now its original pressure *plus* the increase in pressure:

$$P_{gas} = 760mmHg\ +\ 50\ mmHg = 810\ mmHg$$

Applications of the Ideal Gas Equation

Determining the molar mass ($\mathscr{M}$) of a gas or vapor
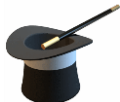
Molar mass ($\mathscr{M}$) is defined as grams of chemical per mole of chemical. We can use the ideal gas law (PV=nRT) and the expression for the density of a gas (D=M/V) to find its molar mass

Derivation:

1. Rearrange the ideal gas equation to make moles (n) the subject.

2. Rearrange the density equation to make mass the subject.

3. Divide equation 2 (grams) by equation 1 (moles) to find Molar mass.

Note: Equation 3 can be rearranged to make gas density the subject. This form of the equation is also used in problems (see appendix)

Example: What is the density of $CCl_4(g)$ at 714 Torr and 125°C?

Answer: 4.43g/L (4.43 x10^{-3} g/mL)

Partial Pressure

Discussion: 'Dry' air is composed of 78% N_2, 21 % O_2 and 1% Ar. If the total atmospheric pressure exerted by the 'dry' air is 760 mmHg, then what fraction of this pressure is due to each component?

$pN_2 =$

$pO_2 =$

$pAr =$

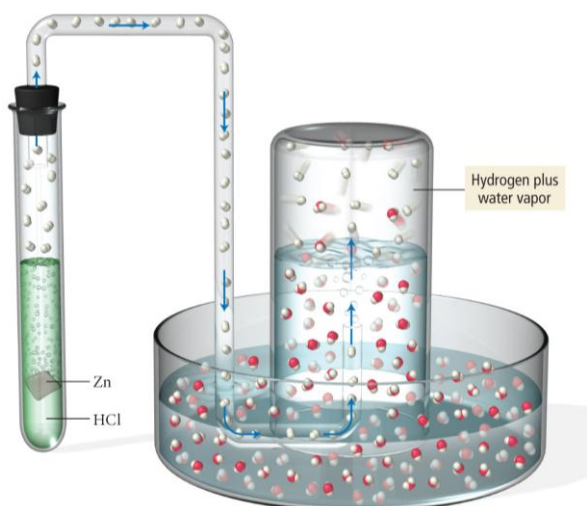|	The total pressure (P_{total}) exerted by a gas mixture = sum of each component's partial pressures ($p_1 + p_2 + p_3$ ……..):

$$P_{total} = p_1 + p_2 + p_3 \text{ ……..}$$

For air: P_{air} = pN_2 + pO_2 + pAr

   i.e   760mmHg   =   593 mmHg   + 160 mmHg   +   7 mmHg

Collecting gases 'over water': In the lab, gaseous reaction products are most often collected 'over water' (recall your 'Determination of R' lab)

Because the gas being collected has been passed through water, it 'drags' some water vapor ('humidity') with it. For the collection of a gas over water:

$$P_{wet\ gas} = p_{dry\ gas} + p_{water\ vapor}$$

Just like on a humid summer's day, the amount (partial pressure) of H_2O vapor (gas) in the air is related to the ambient temperature:

Temperature (°C)	Pressure (mmHg)	Temperature (°C)	Pressure (mmHg)
0	4.58	55	118.2
5	6.54	60	149.6
10	9.21	65	187.5
15	12.79	70	233.7
20	17.55	75	289.1
25	23.78	80	355.1
30	31.86	85	433.6
35	42.23	90	525.8
40	55.40	95	633.9
45	71.97	100	760.0
50	92.6		

If the partial pressure of any components of a gas mixture (including 'wet' gases) are known, this information may be used in ideal gas law (PV = nRT) problems. *Mole fractions* may also be determined.

Example: During an experiment, $H_2(g)$ was collected 'over water'. The *wet* gas recovered was found to have a pressure of 745 Torr at $25^{\circ}C$.

1. What was the partial pressure of the dry H_2 gas?

2. If the container the gas was collected in had a volume of 1.50 L, then how many moles of *dry* gas were collected?

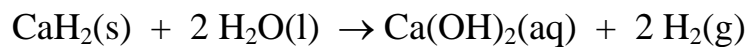3. What is the *mole fraction* of H_2 (g) in the wet gas? What is the mole fraction of H_2O (g)?

Since $P_{comp.} \propto n_{comp.}$ for a mixture of gases (with V and T fixed), then:

$$\text{Mole fraction}_{comp.} = \frac{n_{comp.}}{n_{mixture}} = \frac{P_{comp.}}{P_{mixture}}$$
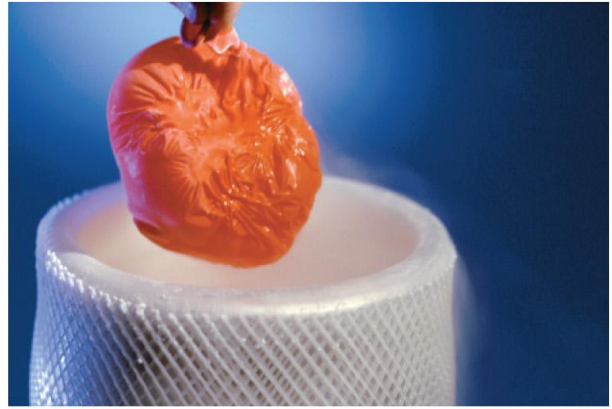
The following reaction between calcium hydride and water is used to inflate life rafts and weather balloons:

$$CaH_2(s) + 2 H_2O(l) \rightarrow Ca(OH)_2(aq) + 2 H_2(g)$$

If 47.0 grams of $CaH_2(s)$ is completely reacted with an excess of water, then:

A. <u>How many moles of hydrogen gas will be produced?</u>

B. <u>What volume (in L) would the hydrogen gas generated in part (a) occupy at $15.0^{\circ}C$ and 725 Torr?</u>

Lower pressure   Higher pressure

Ideal Gas Law

$$PV = nRT$$

| constant n and T | constant n and P | constant P and T |

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

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

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

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

$$V \propto T$$

$$V \propto n$$

Boyle's Law

Charles's Law

Avogadro's Law

Density

$$\underbrace{\frac{n}{V} \quad \mathcal{M}}_{} = \frac{P\mathcal{M}}{RT}$$

Molar density   **Molar mass**

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

$$\frac{\text{moles}}{\text{liter}} \times \frac{\text{grams}}{\text{mole}} = \frac{\text{grams}}{\text{liter}}$$

Molar density   Molar mass   Density in grams/liter