

## Announcements – 10/4/00

- *Quiz and Demos on Friday*
- **Problem Set Solutions**  
- *Now ONLINE!*

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## Charles' Law Example

- If a gas occupies **1.00 L** at **25.0 °C** and, after heating, expands to a volume of **2.00 L**, to what *temperature* was the gas heated?

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

$$T_2 = \frac{(25.0 + 273.15)(2.00 \text{ L})}{(1.00 \text{ L})} = \mathbf{596.3 \text{ K}}$$

$$T_2 = 596.3 \text{ K} - 273.15 = 323.15 \text{ }^\circ\text{C}$$
$$= \mathbf{323. \text{ }^\circ\text{C}}$$

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## Avogadro's Law

- Recall: *Equal volumes of gases (at the same temperature and pressure) contain equal number of molecules.*

So, it follows that:

Volume  $\propto$  # mol gas  
(at constant temperature and pressure)

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## Ideal Gas Law

- Puts together all of the three previous laws:

Boyle: P-V @ constant T and n

Charles: V-T @ constant P and n

Avogadro: V-n @ constant P and T

So: **PV = nRT** Ideal Gas Law

**R** = universal gas constant = **0.082058 L-atm-mol<sup>-1</sup>-K<sup>-1</sup>**

(In S.I. Units: R = **8.3145 J-mol<sup>-1</sup>-K<sup>-1</sup>**)

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## Standard Temp and Pressure

- **One implication of IGL:** *at a particular T and P, a fixed amount of ANY GAS will occupy the same volume*

-Define a **reference condition:**

$$T = 0^{\circ}\text{C} (=273.15 \text{ K})$$

$$P = 1 \text{ atm}$$

Standard Temp and  
Pressure:  
**STP**

Solving IGL for VOLUME:  $V = nRT/P$

$$= \frac{(1 \text{ mol})(0.08206)(273.15 \text{ K})}{1 \text{ atm}}$$

$$= \underline{\underline{22.41 \text{ L}}}$$

(volume occupied by 1 mol of ANY GAS at STP)

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## Tips for using the IGL

- **There are FOUR variables!**  
-you've gotta know THREE to find the 4th one

- **UNITS! UNITS! UNITS!**

-convert all quantities to appropriate units

$$V = L, P = atm, T = K, n = mol$$

-use dimensional analysis to check your units

-HINT: look at the units of **R** as a final check:

$$(0.08206 \text{ L-atm/mol-K})$$

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## Example: P, V and T changes

- For a *fixed amount* (n mol) of gas:

$$P_1V_1 = nRT_1 \text{ - initial condition}$$

$$P_2V_2 = nRT_2 \text{ - final condition}$$

So: 
$$\frac{P_1V_1}{T_1} = nR = \frac{P_2V_2}{T_2}$$

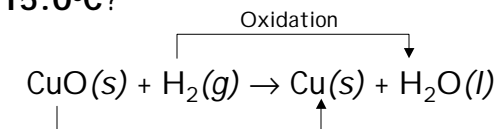
Leaving:

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

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## How does this relate to Chemistry?

- What *mass* of Cu will be produced by the reduction of excess CuO by **250. mL H<sub>2</sub>** at **1.025 atm** and **15.0°C**?



↓ Reduction  
↑

Excess    250. mL    **? grams**  
 1.025 atm  
 15.0 °C

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## The Solution: Gas Stoichiometry

Strategy:

vol/press/temp @ mol H<sub>2</sub> @ mol Cu @ g Cu

**PV = nRT:** solve for n

$$n = \frac{PV}{RT} = \frac{(1.025 \text{ atm})(0.250 \text{ L})}{(0.08206 \text{ L-atm/mol-K})(288.15\text{K})} = 0.010837 \text{ mol H}_2$$

$$0.010837 \text{ mol H}_2 \times \frac{1 \text{ mol Cu}}{1 \text{ mol H}_2} \times \frac{63.546 \text{ g Cu}}{1 \text{ mol Cu}} = 0.68855 \text{ g Cu}$$

$$= \underline{\underline{0.689 \text{ g Cu}}}$$

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## Gas Density

- Density (d) = mass/volume

- Rearranging IGL:

$$\frac{n}{V} = \frac{P}{RT}$$

- Convert mol → grams (mass):

$$\frac{MW \times n}{V} = \frac{P \times MW}{RT}$$

- Now we have density:

$$d = \frac{P(MW)}{RT} = \text{g/L}$$

### Gas Density:

- varies with PRESSURE
- varies with TEMPERATURE
- varies with MOLAR MASS

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## Molar Mass Determination

- If a sample of an unknown gas weighs *0.893 grams* and occupies a volume of *250. mL at STP*, what is its **molar mass** (MW)?

**MW (molar mass) = g/mol**

-solve IGL for mol (n):  $n = \frac{PV}{RT} = \frac{(1 \text{ atm})(0.250 \text{ L})}{(0.08206)(273.15 \text{ K})}$   
 $n = 0.0111534 \text{ mol}$

So:  $MW = 0.893 \text{ g} / 0.0111534 \text{ mol} = 80.065 \text{ g/mol}$   
 $= \underline{\underline{80.1 \text{ g/mol}}}$

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## Gas Mixtures

- Dalton's Law of Partial Pressures:

*The **TOTAL PRESSURE** of a gas mixture is the **SUM of the pressures** exerted by each of the gases in the mixture.*

$$P_{\text{total}} = P_1 + P_2 + P_3 + \dots$$

At constant Temp and Volume:

$$P_1 = n_1 (RT/V) \rightarrow P_1 \propto n_1$$

- partial pressure of a gas in a mixture is proportional to the number of mol of that gas in the mixture

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