

# **Chemistry 102 Winter 2010**

## **Announcements**

1. First midterm exam on Jan. 25, Monday over Chapter 9 and Chapter 10.1-10.5 and 10.7.  
Arrive early to find your assigned seat.  
Bring a calculator.

## **Today**

1. Continue with behavior of gases.

**Ideal gas law**

$$PV = nRT$$

R = universal gas constant

$$R = 0.08206 \frac{\text{L-atm}}{\text{mol-K}}$$

$$R = 8.314 \frac{\text{J}}{\text{mol-K}}$$

Practice: In the reaction of oxygen and hydrogen to make water, how many liters of hydrogen at 0 °C and 1 atm are needed to react with 10.0 g of oxygen?

## Combined gas law

V decreases with pressure ( $V \propto 1/P$ ).

V increases with temperature ( $V \propto T$ ).

V increases with particles ( $V \propto n$ ).

Combining these ideas:

$$V \propto \frac{n \times T}{P}$$

For a given amount of gas (constant number of particles,  $n_{\text{initial}} = n_{\text{after}}$ ) that changes in volume, temperature, or pressure:

$$V_{\text{initial}} \propto \frac{n_{\text{initial}} \times T_{\text{initial}}}{P_{\text{initial}}}$$

$$V_{\text{after}} \propto \frac{n_{\text{after}} \times T_{\text{after}}}{P_{\text{after}}}$$

$$n_{\text{initial}} = n_{\text{after}}$$

$$\frac{V_{\text{initial}} \times P_{\text{initial}}}{T_{\text{initial}}} = \frac{V_{\text{after}} \times P_{\text{after}}}{T_{\text{after}}}$$

Example: A sample of gas is pressurized to 1.80 atm and volume of 90.0 cm<sup>3</sup>. What volume would the gas have if it were pressurized to 4.00 atm without a change in temperature?

Example: A 2547-mL sample of gas is at 421 °C. What is the new gas temperature if the gas is cooled at constant pressure until it has a volume of 1616 mL?

Practice: A balloon is filled with helium. At 26 °C, its volume is 5.90 L. If the balloon's temperature increases to 278 °C at constant pressure, what is the final volume of the balloon?

Practice: A 2.25-L balloon has internal pressure of 1.10 atm. If the balloon increases to 7.05 L at constant temperature, what is the final pressure inside the balloon?

## Gas mixtures and partial pressures

Fig. 10.13, p. 361.

For a mixture of nitrogen and oxygen at the same temperature and volume,

$$P_{\text{mixture}} = P_{\text{N}_2} + P_{\text{O}_2} = P_{\text{total}}$$

Each component in the mixture contributes partially to the overall pressure. The partial pressure contribution ( $P_{\text{N}_2}$ ) is based on the number of moles of that component ( $n_{\text{N}_2}$ ).

$$PV = nRT$$

$$P_{\text{N}_2} = n_{\text{N}_2}RT/V \quad \text{and} \quad P_{\text{O}_2} = n_{\text{O}_2}RT/V$$

The total number of moles ( $n_{\text{total}}$ ) is responsible for the overall pressure of a mixture ( $P_{\text{total}}$ ):

$$P_{\text{total}} = n_1RT/V + n_2RT/V + n_3RT/V + \dots$$

$$P_{\text{total}} = (n_1 + n_2 + n_3 + \dots) \times RT/V = n_{\text{total}}RT/V$$

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



Example: Given 2.0 mol O<sub>2</sub> and 8.0 mol N<sub>2</sub> inside a 3.0-L container at 273 K, calculate the partial pressure from O<sub>2</sub>, the partial pressure from N<sub>2</sub>, and the total pressure in units of atmospheres. R = 0.08206 L-atm/mol-K

mole fraction: the ratio of the number of moles of one component to the total number of moles in a mixture of components.

In the previous example, what is the mole fraction of  $\text{N}_2$ ?  $\text{O}_2$ ?

Water vapor can be a component. Table 10.4.

Example: P-S Practice 10.11, p. 364.

## Real gas behavior

$P_{\text{ideal}} V_{\text{ideal}} = nRT$  holds for ideal gas behavior.

Ideal gas behavior is based on KMT, p. 345:

1. Size of particles is small compared to the distances between particles. (The volume of the particles themselves is ignored)

So a real gas actually has a larger volume than an ideal gas.

$$V_{\text{real}} = V_{\text{ideal}} + nb \quad n = \text{number (in moles) of particles}$$

$$V_{\text{ideal}} = V_{\text{real}} - nb \quad b = \text{van der Waals constant } b \text{ (Table 10.5)}$$

$$P_{\text{ideal}} (V_{\text{real}} - nb) = nRT$$

But this only matters for gases in a small volume and with a high pressure because gas molecules are crowded and bumping into each other.

Ideal gas behavior is based on KMT, p. 345:

3. Particles do not interact, except during collisions.

Fig. 10.16, p. 365

Particles in reality attract each other due to dispersion (London) forces. They “tug” on each other and soften the impact of collisions.

So a real gas actually has a smaller pressure than an ideal gas.

$$P_{\text{real}} = P_{\text{ideal}} - \frac{n^2 a}{V_{\text{real}}^2} \quad n = \text{number (in moles) of particles}$$

$$P_{\text{ideal}} = P_{\text{real}} + \frac{n^2 a}{V_{\text{real}}^2} \quad a = \text{van der Waals constant } a \text{ (Table 10.5)}$$

$$(P_{\text{real}} + n^2 a / V_{\text{real}}^2) (V_{\text{real}} - nb) = nRT$$

Again, only matters for gases in a small volume and with a high pressure because gas molecules are crowded and bumping into each other.

## Gas density and molar masses

We can re-arrange the ideal gas law and use it to determine the density of a gas, based on the physical properties of the gas.

$$PV = nRT$$

$$\text{molar mass} = MM = M = \text{mass/moles} = m/n \\ \Rightarrow n = m/MM$$

$$PV = (m/MM)RT$$

$$m/V = P * MM/RT$$

$$\text{density} = d = \text{mass/volume} = m/V = P*MM/RT$$

$$m/V = P*MM/RT$$

If we don't know the identity of a gas (don't know its MM), then we can measure its mass, volume, pressure, and temperature to solve for its MM – and its probable identity.

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**OWL HW deadlines**

Due Date	Assignment
1/27/10	9 – Molecular Structures
1/27/10	10 – Gases and the Atmosphere
2/10/10	11 – Liquids, Solids, and Materials
2/24/10	12 – Chemical Kinetics
3/15/10	13 – Chemical Equilibrium

**Exam review**

1. Lecture notes and examples.
2. OWL HW.
3. Recitation HW and quizzes.
4. Extra Chapter 9 Problems: 1-12, 22-29, 31-34, 37-39, 43-52.
5. Extra Chapter 10 Problems: 16, 20-43, 68-80.

**Before next class,**

1. Study for exam.
2. Work on OWL HW.