Department of Chemistry
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Kinetic Molecular Theory – Supplemental Worksheet

1. What assumptions do we make when using the ideal gas laws? Are the assumptions always true?

Gas molecules are infinitely small
Particles are in constant motion
Particles do not exert forces on each other
Average kinetic energy is directly proportional to Temperature.
The 1st and 3rd assumptions listed fail at higher pressures.

2. Consider 2 gases, A and B, each in a 1.0 L container with both gases at the same temperature and pressure. The mass of gas A in the container is 0.25 g and the mass of gas B in the container is 0.51 g.
   A. Which gas sample has the most molecules present?
   B. Which gas sample has the largest average kinetic energy?
   C. Which gas sample has the fastest average velocity?
   D. How can the pressure in the 2 containers be equal to each other since the larger gas B molecules collide with the container walls more forcefully?

V, T, & P are all constant, so n must be constant as well. Equal numbers of moles means gas B molecules must be heavier than gas A molecules.

A. n is constant, same number of molecules.
B. T is constant, KE_{avg} is the same ($KE_{avg} = \frac{3}{2}RT$)
C. A is light and will have a faster average velocity
D. The heavier gas B molecules collide more forcefully, but gas A molecules have a faster average velocity and collide more frequently.

3. Calculate the average kinetic energies of CH_4 and N_2 molecules at 5°C and 112°C.

At each temperature CH_4 & N_2 will have the same average KE. $KE = \frac{3}{2}RT$

$$\text{At 5°C, } KE = \frac{3}{2}\times\left(\frac{8.314\ J}{K\ mol}\right)\times 278K = 3.47\times10^{3}\ \frac{J}{mol}$$

To determine the energy per molecules, divide by Avogadro's number for molecules: 6.022×10^{23} molecules/mole.

$$\frac{3.47\times10^{3}\ \frac{J}{mol}}{\frac{6.022\times10^{23}\ molecules}{mole}} = 5.76\times10^{-21}\ \frac{J}{molecules}$$

$$\text{At 112°C, } KE = \frac{3}{2}\times\left(\frac{8.314\ J}{K\ mol}\right)\times 385K = \frac{4.80\times10^{3}\frac{J}{mol}}{\frac{6.022\times10^{23}\ molecules}{mole}} = 7.97\times10^{-21}\ \frac{J}{molecules}$$

4. Calculate the root mean square for nitrogen (g) at 217°C, helium (g) at 75°C, and xenon (g) at 27°C.

$$V_{rms} = \left(\frac{3RT}{M}\right)^{1/2}$$

N_2: $V_{rms} = \left(\frac{3 \times \frac{8.314\, kg\, m^2}{s^2\, K\, mol} \times 490\, K}{\frac{28.02\times 10^{-3} kg}{mol}}\right)^{1/2} = \frac{660\, m}{s}$

He : $V_{rms} = \left(\frac{3 \times \frac{8.314\, kg\, m^2}{s^2\, K\, mol} \times 348\, K}{\frac{4\times 10^{-3} kg}{mol}}\right)^{1/2} = \frac{1473\, m}{s}$

Xe: $V_{rms} = \left(\frac{3 \times \frac{8.314\, kg\, m^2}{s^2\, K\, mol} \times 300\, K}{\frac{131.3\times 10^{-3} kg}{mol}}\right)^{1/2} = \frac{238\, m}{s}$

5. A 100 L flask contains a mixture of methane and argon at 27°C. The mass of the argon present is 245 g and the mole fraction of methane in the mixture is 0.623. Calculate the total kinetic energy of the gaseous mixture.

$n_{Ar} = \frac{245g}{\frac{39.95g}{mol}} = 6.13\ mol\ Ar$; $\chi_{CH_4} = \frac{n_{CH_4}}{n_{CH_4} + n_{Ar}}$ $0.623 = \frac{n_{CH_4}}{n_{CH_4} + 6.13}$ $n_{CH_4} = 10.1\ mol\ CH_4$

$KE_{avg} = \frac{3}{2} RT\ for\ 1\ mol$ $KE_{total} = (10.1 + 6.13 mol) \times \frac{3}{2} \times \frac{8.314\, J}{K\, mol} \times 300K = 6.07\times 10^4 J = 6.07kJ$

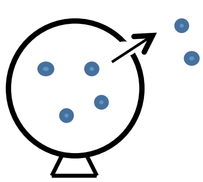
6. True or False in regards to the Maxwell-Boltzmann distribution

A. The distributions are symmetrical
B. The more massive the particle the faster the velocities.
C. The lower the temperature the slower the average velocity.
D. Broader distributions are caused by higher temperatures and heavier particles.

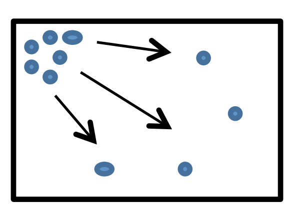
A. False: Asymmetrical
B. False: The lighter the particle the faster the velocity
C. True: velocity is temperature dependent
D. False: Higher temperatures do cause a broader distribution, but lighter particles cause broader distributions not heavier particles.

7. Draw examples and explain effusion and diffusion.

Effusion: rate at which a gas escapes Through a small hole.



Diffusion: rate at which a gas travels a cross a room.



Note for problems #8 & 9

Graham's law of effusion states that the rate is inversely proportional to the square root of the mass of its particles.

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

8. Freon-12 is used as a refrigerant in central home air conditioners. The rate of effusion of Freon-12 to Freon-11 (molar mass = 137.4g/mol) is 1.07:1. The formula of Freon-12 is one of the following: CF_4, CF_3Cl, CF_2Cl_2, $CFCL_3$, or CCl_4. Which formula is correct for Freon-12?

Let Freon-12 = gas 1 and Freon-11 = gas 2:

$$\frac{1.07}{1.00} = \sqrt{\frac{137.4}{M_1}} \quad , \quad 1.14 = \frac{137.4}{M_1} \quad , \quad M_1 = 121 g/mol$$

The molar mass of CF_2Cl_2 is equal to 121g/mol, so Freon -12 is CF_2Cl_2.

9. The rate of effusion of a particular gas was measured to be 27.2 mL/min. under the same conditions, the rate effusion of pure methane gas (CH_4) is 47.8mL/min. What is the molar mass of the unknown gas?

$Rate_1$ = 27.2mL/min $Rate_2$ = 47.8mL/min M_2 = 16.04g/mol M_1 = ?

$$\frac{27.2}{47.8} = \sqrt{\frac{16.04}{M_1}} \quad ; \quad 0.569 = \sqrt{\frac{16.04}{M_1}} \quad ; \quad M_1 = 49.5 g/mol$$