Unit 5 Name \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

Gases Date \_\_\_\_\_\_\_\_\_\_\_\_ Block \_\_\_\_

Unit 5B – Pressure and Gas Laws

### Knowledge/Understanding Goals:

* difference between ideal and “real” gases
* when a gas exhibits ideal vs. non-ideal behavior

### Skills:

* van der Waals’ Gas Law

###  Notes:

ideal gas: a gas that behaves according to KMT. In a phase diagram, gases behave ideally when they are sufficiently far from a phase transition (*i.e.,* sufficiently far from the boundaries of the gas region of the phase diagram).

real gas: a gas that is deviating from KMT in some way (usually at low temperatures and/or high pressures, which would cause it to be close to its condensation or deposition point).

In any chemistry class and on the AP test, *assume that all gases are ideal* (and that the ideal gas law applies) unless you have specific information that would indicate otherwise.



The Dutch physicist Johannes Diderik van der Waals proposed an expanded version of the ideal gas law that compensated for intermolecular effects at temperatures and pressures near the sublimation and vaporization curves. The van der Waals Equation is:



Deep breath…the equation is given on the AP test and you will be provided with all of the other values you need!

* The constant *“a”* is related to the attractions between the particles and must be determined experimentally (ie: would be given to you in a table).
* The constant *“b”*  is related to the volume excluded by a mole of particles and must be determined experimentally (ie: would be given to you in a table).

### The Pressure Correction Term (a)

When the gas is exhibiting ideal behavior, the molecules are exhibiting essentially no attraction to one another per KMT, and the constant *“a”* is essentially zero.

However, at high pressures, the molecules are close enough to one another that they start to form clusters. This effectively decreases the pressure, because a cluster behaves like a single (albeit more massive) molecule, decreasing the number of separate molecules that are present.

Similarly, at extremely low temperatures (near absolute zero) the molecules are moving and colliding so slowly that the kinetic energy of the molecules is no longer enough to overcome other interactions (attractions/repulsions) between the molecules.

Thus in the van der Waals equation, a correction factor (a) is added to the pressure term to account for these behaviors.

### The Volume Correction Term

When the gas is behaving ideally, the volumes of the gas molecules themselves are such a small fraction of the total volume that they can be ignored, and the constant *b* is essentially zero.

However, at extremely high pressures, we can no longer assume that gas molecules fill no volume. Their volumes start to “get in the way” of further compression, and the observed volume is greater than that predicted by PV = nRT.

Thus in the van der Waals equation, a correction factor (b) is subtracted from the volume term to account for these behaviors.

### The Constants “a” and “b”

The constants “a” and “b” are specific to the particular gas, and are determined empirically (experimentally/mathematically). The following are values of the constants “a” and “b” for selected gases, which you would be provided on a test.

|  |
| --- |
| **van der Waals Constants** |
| **Substance** | **a** | **b** |
| acetone | 14.28 | 0.0994 |
| NH3 | 4.281 | 0.03707 |
| CO2 | 3.688 | 0.04267 |
| CCl4 | 20.01 | 0.1281 |
| Cl2 | 6.666 | 0.05622 |
| He | 0.0350 | 0.0237 |
| H2 | 0.2509 | 0.02661 |
| CH4 | 2.313 | 0.04278 |
| SO2 | 6.893 | 0.05636 |
| H2O | 5.609 | 0.03049 |

**Gas Laws Practice Problems**

**Problem #1:** A 1L mixture of nitrogen and neon gases contains equal moles of each gas (0.208 mol) and has a total mass of 10.0 g. What is the density of this gas mixture at 500 K and 15.0 atm? Assume ideal gas behavior.

***density = 8.8 g/L***

**Problem #2:** Three 1.00 L flasks at 25.0 °C and 1 atm pressure contain: CH4 (flask A), CO2 (flask B) and NH3 (flask C). Which flask (or none) contains 0.041 mol of gas?

***All three flasks contain 0.041 mol of the different gases***

**Problem #3:** 1.00 L of liquid nitrogen is kept in a closet measuring 1.00 m by 1.00 m by 2.00 m. (1m3 = 1L) Assuming that the container is completely full, that the temperature is 25.0 °C, and that the atmospheric pressure is 1.00 atm, calculate the percent (by volume) of air that would be displaced if all the liquid nitrogen evaporated. (Liquid nitrogen has a density of 0.807 g/mL.)

***704 L N2, 35.2% of the air gets displaced***

**Problem #4:** A sample of gas (1.90 mol) is in a flask at 21.0 °C and 0.917 atm. The flask is now opened and more gas is added to the flask. The new pressure is 1.05 atm and the temperature is now 26.0 °C. How many moles of gas are now in the flask?

***2.13 mol***

**Problem #5:** 100. g of our atmosphere is a mixture of gases (roughly 79% N2, 20% O2 and 1%Ar).

1. What is the partial pressure (in atm) of each gas in the atmosphere?
2. A mixture of He and O2 gases is used by deep sea divers. If the pressure of the gas a diver inhales is 8.00 atm what percent of the mixture should be O2, if the partial pressure of O2 is to be the same as what the divers would ordinarily breathe at sea level?
3. ***0.813 atm N2, 0.180 atm O2, 0.007 atm Ar b) 15.5% O2***

Solutions

**Solution to 1:**

1. total moles in mixture = 0.416
grams of mixture = 10.0 g
2. solve for volume with PV = nRT (1.14L)
3. solve density

**Solution to 3:**

1) calculate grams, then moles of N2:

 0.807 g mL¯1 x 1000 mL = 807 g

 807 g / 28.014 g mol¯1 = 28.8070 mol

2) Calculate volume of N2 at stated pressure and temperature:

 V = nRT / P

 V = [(28.8070) (0.08206) (298)] / 1.00

 V = 704.44 L

3) Calculate volume of closet in liters:

 1.0 m x 1.0 m x 2.0 m = 2.0 m3

 2.0 m3 = 2000 L

4) Assume N2 displaces 704.44 L of air:

 704.44 L/ 2000 L = 0.3522

 35.2% of the air gets displaced.

**Solution to 4:**

1) Use PV = nRT with the first set of data to get the volume of the container:

(697.0/760.0) (x) = (1.90 mol) (0.08206 L atm/mol K) (294.0 K)

x = 49.9819572 L

2) Use PV = nRT with the second set of data, using the volume just calculated. Solve for moles:

(795.0/760.0) (49.9819572 L) = (x) (0.08206 L atm/mol K) (299.0 K)

x = 2.13 mol

**Solution to 5(a):**

1) Determine the mole fraction of each gas. Assume 100 g of atmosphere:

N2: 79 g / 28.0 g mol¯1 = 2.82 mol
O2: 20 g / 32.0 g mol¯1 = 0.625 mol
Ar: 1 g / 40 g mol¯1 = 0.025 mol

2) Determine mole fraction of oxygen:

0.625 mol / 3.47 mol = 0.180

3) Determine partial pressure of oxygen:

1.00 atm x 0.180 = 0.180 atm

**Solution to 5(b):**

1) Calculate mole fraction of He/O2 mixture:

0.180 / 8 = 0.0225 mol of O2
7.820 / 8 = 0.9775 mol of He

2) Convert each to grams:

O2: 0.0225 mol x 32.0 g mol¯1 = 0.72 g
He: 0.9775 mol x 4.00 g mol¯1 = 3.91 g

3) Calculate percent of O2 in the mixture:

0.72 g / 4.63 g = 0.1555 = 15.55%