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

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

Unit 5A – Pressure and Gas Laws

### Knowledge/Understanding Goals:

* recall gas laws from Chemistry
* what partial pressure means

### Skills:

* pressure conversion
* ideal & combined gas law problems
* solve problems using a combination of Dalton’s Law, Raoult’s Law, and the Ideal Gas Law

## Notes:

**Kinetic Molecular Theory (KMT)**

The Kinetic-Molecular Theory of gases (KMT) states that gases are made of (statistically) extremely large numbers of molecules that:

1. Are in constant motion
2. Are, with respect to their diameter, far apart
3. Collide elastically with each other and the walls of the container they are housed in

Interactions between the molecules (other than collisions), such as attraction (phase changes) or chemical reactions, are considered separately from KMT.

pressure (P): the force exerted by the molecules, due to collisions with the walls of the container, per unit of surface area of the container (units = atm, bar, mmHg)

volume (V): the space the molecules take up (units = L)

moles (n): a measure of the number of molecules

temperature (T): a measure of the average kinetic energy of the molecules. (units = K)

Recall that KE = **½mv2**

where m = mass (of a molecule) and v = velocity (of that molecule).

Essentially, what this shows is when the temperature of a substance is increased, the velocity of the molecule increases, assuming no chemical reaction occurred to change the mass.

## Pressure & Pressure Conversions

### Pressure Conversions:

Chemists prefer to work in atmospheres. Physicists tend to prefer S.I. units, such as kPa.

1 atm = 101.325 kPa = 1.01325 bar

1 atm = 760 mm Hg = 760 torr

1 atm = 14.696  = 14.696 p.s.i.

\*Unit conversions are provided on the AP test for pressure between atm, mmHg, and torr.

\*\*Watch your units! The unit value for pressure will determine which R constant you use in the gas law equations.

Gas constant **R =**  8.31 J **/** mol∙K

0.0821 Latm **/** mol∙K

62.4 L torr **/** mol∙K

### Manometers & Barometers

There are many ways of measuring pressure. For the AP exam, you need to be able to determine pressure from manometer and barometer readings.

manometer: a device that measures pressure by measuring a difference in the level of a liquid. The “classic” manometer is a U-shaped tube, with a sample of gas (whose pressure you want to measure) on one side. The other side is either open to the atmosphere (open manometer), or is sealed with a total vacuum on the other side (closed manometer).

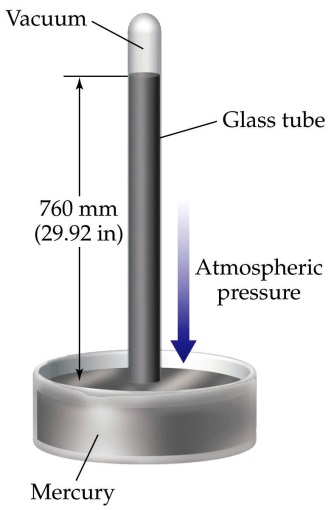
Although mercury manometers are uncommon because of the toxicity of mercury, any manometer problem that you are likely to see will involve a mercury-filled manometer. In fact, the pressure unit “mm Hg” comes from the height (or difference in height) of a column of mercury in a manometer.

The principle at work here is that your mercury column opposite of your gas sample will move up or down depending upon the pressure difference between the atmospheric pressure in one arm of the U-tube and the sample pressure in the bulb end of the U-tube.

If the column moves up from the original position: sample pressure > atmospheric pressure

If the column moves down from the original position: atmospheric pressure > sample pressure

|  |  |
| --- | --- |
|  |  |
| open manometer | closed manometer |

Barometers follow a similar principle of comparing the pressure exerted by a sample of gas versus the atmospheric pressure of the room. Atmospheric pressure is exerted on the mercury in the dish, forcing liquid up into the barometer, while sample pressure pushed the liquid level down the barometer. If no gas is in the barometer, it will read the atmospheric pressure.

**Ideal Gas Law**

Ideal gases follow the formula:

*P V* = *n*R*T*

P = pressure n = (number of) moles

V = volume R = gas constant

T = temperature (Kelvin)

* If you are given only one set of conditions for a gas for a gas (nothing is changing), use this formula.
* Choose a value of R that has exactly the same units as the numbers you are given in the problem. If this is not possible, you must convert the units from the problem to units that agree with your R value.



\*Again, the AP test provides conversion values for R (atm, torr, and joule)

* Temperatures must be in Kelvin. (K = C + 273)\* \*Also provided on the test
  + STP = Standard Temperature and Pressure
    - Temp = 273 K
    - Pressure = 1 atm
  + Under STP: the Ideal Gas Law shows that 1 mol of gas will occupy what volume?
* Remember that you can calculate moles (n) from the mass of the gas (in grams) divided by its molar mass .

**Practice Problems**

1. What pressure would a 2.3 mole sample of oxygen gas exert on a 5000.0 mL scuba tank when it acclimates to the same temperature as 15⁰C sea water?
2. How many grams of carbon dioxide are left in a 2.5 liter fire extinguisher if the pressure gauge reads 1345 kPa at room temperature (25⁰C)?

**Make sure you understand the relationship between these variables and WHY they change when one variable is altered.**

## Ex: What is / what causes pressure? Draw a diagram to support your answer.

## Ex 2: If volume is decreased, how would pressure change and why?

## Ex 3: How would two containers behave during a temperature increase if one has rigid walls while the other one has elastic walls? Explain.

## Ex 4: How would adding moles of gas to a container alter the container if volume is held constant (rigid walls)?

**Combined Gas Law**

Because the ideal gas law is always true (unless the gas is behaving non-ideally), *PV* = *n*R*T* at time 1 and at time 2. This means:



\*Since R is a constant value, both equations can be set equal to one another

This is an expanded version of the combined gas law, which is usually expressed as:



* Use this law when the gas is changing from one set of conditions to another.
* Cancel (*i.e.* drop) any variable that is not changing. (Assume anything not explicitly mentioned in the problem is not changing.)
* Units must be the same on both sides of the equal sign
* Temperatures must be Kelvin.

## Notes: - A container with rigid walls = volume is held constant

## - A container with elastic walls = pressure is held constant

## - Unless specifically mentioned, assume a variable is being held constant (especially moles of gas)

**Practice Problems**

1. A sample of air has a volume of 550.0mL at 106oC. At what temperature will its volume be 700.0mL at constant pressure?
2. A sample of gas at 104oC and 0.870 atm occupies a volume of 3.0L. What volume would this gas occupy at 60oC and 1.7 atm?
3. A sample of carbon dioxide contained in a 2.4L case has a pressure of 1.3 atm at standard temperature. How many grams of gas are in the container?

The AP board **LOVES** conceptual questions using particulate diagrams; don’t be surprised to see one involving the gas laws:

1. Which of the following particulate diagrams represents a scenario where an increase in temperature is observed in the system?

The ideal and combined gas laws can each be utilized to describe the mole, pressure, temperature, and volume characteristics of a whole sample of gas. However, if that sample of gas is a solution of different molecules, it is often advantageous to be able to describe the effect of each individual species present in the sample.

Partial Pressure: the partial pressure of a gas is the pressure due to only the molecules of that

gas.

Dalton’s Law of Partial Pressures: the sum of all the partial pressures in a sealed container equals the total pressure.

*P* = *P*T = *P*1 + *P*2 + *P*3 + …

(To make things more clear, we will use *P*T to mean the total pressure.)

mole fraction (*χ*): the fraction of the total moles (or molecules that are the compound of interest. For example, if we have 20 moles of gas, and 9 moles are N2, the mole fraction of N2 is:



For example, suppose we had the following tank, with a total pressure of 1.00 atm:



If we ignore all of the molecules except for nitrogen, the tank would look like this:



If 45% of the molecules are nitrogen (*χ*N­2 = 0.45), then the pressure just from these nitrogen molecules (the partial pressure of nitrogen) must be 0.45 times the total pressure of 1 atm. This means:

*P*N­2 = *χ*N­2*P*T

*P*N­2 = (0.45)(1 atm) = 0.45 atm

Similarly, because 55% of the molecules are oxygen, this means:

*χ*O­2 = 0.55

*P*O­2 = *χ*O­2*P*T      *P*O­2 =(0.55)(1 atm) = 0.55 atm

Note that the two partial pressures add up to the total pressure:

*P*T = *P*N­2 + *P*O­2   = 0.45 atm + 0.55 atm  =  1 atm

**Using Dalton’s Law with the Ideal Gas Law**

Recall the two tanks from our example. Assuming N2 and O2 are behaving like ideal gases, the ideal gas law must be true in both tanks.

|  |  |
| --- | --- |
|  |  |
| *P*= *P*T        *n* = *n*T  *P*T*V* = *n* TR*T* | *P*= *P*N­2       *n* = *n*N­2  *P*N­2*V* = *n* N­2R*T* |

*In other words, the ideal gas law can be used either with the total moles and total pressure, or with the moles of one specific gas and the partial pressure of that gas.*

### Sample problem:

A 12.0 L tank of gas has a temperature of 30.0°C and a total pressure of 1.75 atm. If the partial pressure of oxygen in the tank is 0.350 atm, how many moles of oxygen are in the tank? How many total moles of gas are in the tank?

**Solution:**

For oxygen: Therefore:

*P*O­2 = 0.350 atm *P*O­2*V* = *n* O­2R*T*

V = 12.0 L (0.350)(12.0) = *n* O­2 (0.0821)(303.0)

n = n *n* O­2 = 0.169 mol

R = 

T = 30.0°C + 273 = 303.0 K

You could figure out the total moles two ways. One is to use the ideal gas law on the total moles:

For total moles: Therefore:

*P* = 1.75 atm *PV* = *n*R*T*

V = 12.0 L (1.75)(12.0) = *n*(0.0821)(303.0)

n = n *n* = 0.844 mol

R = 

T = 30.0°C + 273 = 303.0 K

The other way to find the total moles is to use the mole fraction and the partial pressure:

*P*O­2 = *χ*O­2*P*T

We know that … Therefore…

*P*O­2 = 0.350 atm 0.350 atm = *χ*O­2(1.75 atm)

*P*T = 1.75 atm



Now that we know the mole fraction of O2, we can figure out the total moles:

