**Answer Key to Gas Laws Problem Set**

The first four problems use the ideal gas law. Notice that pressure, volume, temperature and moles are not changing. When using the ideal gas law, PV = nRT, the constant, R = 0.08206 L∙atm/mol∙K, dictates the units of pressure, volume and temperature. The pressure unit must be in atmospheres, the volume unit must be in liters, and the temperature unit must be in Kelvin.

**1. What is the pressure in atmospheres of 0.0233 mol an ideal gas at 27°C in a 52.9 mL container?**

Let’s collect the terms we need and make sure they are in the proper units:

 n = 0.0233 mol (no conversion is necessary here)

 T = 27°C + 273 = 300K (temperature MUST be in Kelvin)

 R = 0.08206 L∙atm/mol∙K

 Rearrange PV = nRT and solve for pressure by dividing through by V:

**2. What is the volume in liters of 2.33 mol of an ideal gas at 27°C at a pressure of 872 torr?**

Let’s collect the terms we need and make sure they are in the proper units:

 n = 2.33 mol (no conversion necessary here)

 T = 27°C + 273 = 300K (temperature MUST be in Kelvin, as stated above)

 R = 0.08206 L∙atm/mol∙K

 Rearrange PV = nRT and solve for volume by dividing through be P:

**3. What is the temperature in Kelvin of 0.0233 mol of an ideal gas at a pressure of 458 torr and a volume of 1870 mL?**

Let’s collect the terms we need and make sure they are in the proper units:

 n = 0.0233 mol (no conversion necessary here)

3. (continued)

 R = 0.08206 L∙atm/mol∙K

Rearrange PV = nRT and solve for temperature by dividing through by nR:

**4. How many moles of ideal gas are present in a 347 mL container with a pressure of 3.88 atmospheres at a temperature of 383 °C?**

Let’s collect the terms we need and make sure they are in the proper units:

 P = 3.88 atm (no converstion is needed)

 T = 383°C + 273 = 656 K (temperature MUST be in Kelvin)

Rearrange PV = nRT and solve for moles by dividing by RT:

The next series of problems deal with the various gas laws in which pressure, volume, moles and or temperature conditions change. One way to determine which gas law to use is to consider a gas law you can name after yourself, in which only R is constant:

 (Insert you name here!) Law:

 (I’ll call it Ladon’s Law, after myself for the purpose of this tutorial)

As you read the problem, cross out any variable not being used to leave behind the equation you need. You’ll see how this applies in the problems and solutions that follow.

These equations are ratios, thus pressure and volume units may be any unit you choose as long as they are consistent. However, temperature still must be converted to Kelvin. The temperature cannot be negative, or have a value of zero as these will lead to pressures and volumes that are impossible. Pressure and volume can decrease to a value of zero, but they cannot be negative. If you are wondering about 0K, a gas does not exist at absolute zero, so that will not be an issue.

**5. An ideal gas initially at a pressure of 0.233 atm in a 5.32 L piston is compressed to a volume of 3.88 L. What is the new pressure of the gas in atmospheres and in torr?**

Notice that only pressure and volume are changing in this problem. Cross out moles (n) and temperature (T) in Ladon’s Law to reveal the equation you need to solve the problem:

The gas law you need is P1V1 = P2V2 , known as Boyle’s Law. Assign pressures and volumes to the variables:

 P1 = 0.233 atm P2 = ? atm

 V1 = 5.32 L V2 = 3.88 L

Solve for the unknown variable, P2, by dividing through by V2:

Notice how the liter units cancel to leave pressure units in atm.

**6. An ideal gas initially at a pressure of 459 torr in a 1438 mL piston has its pressure adjusted to 1.79 atm. What is the new volume of the piston?**

Notice that only pressure and volume are changing in this problem. Cross out moles (n) and temperature (T) in Ladon’s Law to reveal the equation you need to solve the problem:

The gas law you need is P1V1 = P2V2 , known as Boyle’s Law. Assign pressures and volumes to the variables:

 P1 = 459 torr P2 = 1.79 atm

 V1 = 1438 mL V2 = ? mL

Notice the pressure units are not consistent. Let’s convert P2 = 1.79 atm to torr:

Rearrange the equation to solve for V2 by dividing through by P2:

Notice that we did not need to have volume in liters and pressure in atmospheres. This is because the equation is a ratio, and the gas law constant, R, is not involved. The torr units cancel to leave volume units of milliliters.

**7. An ideal gas initially at a temperature of 18.44°C in a 5.00 L container is heated to 127.33°C. What is the new volume of the container in L?**

Notice that only volume and temperature are changing. Cross out moles and pressure in Ladon’s Law to reveal the equation you need to solve the problem:

The gas law required is Charles’ Law:

Assign the volume and temperatures to the variables. You must convert the temperatures to Kelvin:

 V1 = 5.00 L T1 = 18.44°C + 273.15 = 291.59 K

 V2 = ? L T2 = 127.33°C + 273.15 = 400.48 K

Rearrange Charles’ Law to solve for V2 by multiplying through by T2:

Notice how the Kelvin units cancel out to leave the volume in L. Convert these volume units the milliliter units asked for in the problem:

**8. An ideal gas initially at a temperature of 345°C and a volume of 478mL is compressed to a volume of 233 mL. What is the new temperature of the gas in Celsius?**

Notice that only volume and temperature are changing. Cross out moles and pressure in Ladon’s Law to reveal the equation you need to solve the problem:

The gas law required is Charles’ Law:

Assign the volume and temperatures to the variables. You must convert the temperatures to Kelvin:

 V1 = 478 mL T1 = 345°C + 273 = 618 K

 V2 = 233 mL T2 = ?

8. (continued)

Rearrange Charles’ Law to solve for T2. One easy way is to cross-multiply and divide through by V1:

Dividing through by V1 yields:

Notice how the volume units in milliliters cancel to leave temperature units of Kelvin. We now need to convert the Kelvin units back into Celsius, as asked for in the problem.

 T2 = 301 K – 273 = 28°C

**9. An ideal gas initially at a temperature of 445K at a pressure of 553 atmospheres is cooled to a temperature of 37°C in a rigid container. What is the new pressure of the gas in the container in atmosphere and torr units?**

Notice that only temperature and pressure are changing. Cross out moles and volume in Ladon’s Law to reveal the equation you need to solve the problem:

The gas law required is the Guy-Lussac Law:

Assign the pressure and temperatures to the variables. You must convert the temperatures to Kelvin:

 P1 = 553 atm T1 = 445 K (already in K)

 P2 = ? atm T2 = 37°C + 273 = 310 K

Rearrange the Guy-Lussac Law to solve for the pressure, P2, by multiplying through by the temperature, T2.

Notice how the Kelvin units cancel to leave pressure units of atmospheres.

The pressure in torr would be:

**10. An ideal gas initially at a temperature of 388K and a pressure of 3.87 atmospheres in a 288 mL container is cooled to a temperature of 125K and the pressure is raised to 6.33 atmospheres. What is the new volume of the container in milliliters?**

Notice that pressure, volume and temperature are changing. Cross out the mole term in Ladon’s Law to reveal the equation you need to solve the problem:

The equation required is often referred to as the combined gas law:

Assign the pressures, volumes and temperatures to the variables.

 P1 = 3.87 atm P2 = 6.33 atm

 V1 = 288 mL V2 = ? mL

 T1 = 388 K T2 = 125 K

Rearrange and solve for V2 by multiplying through by T2 and dividing through by P2:

Notice how the pressure of atmospheres and temperature units of Kelvin cancel to leave volume units of milliliters.

**11. If 15.0 moles of an ideal gas has a pressure of 3.55 atmospheres at constant volume and temperature, how many moles of gas will have a pressure of 9.45 atmospheres?**

Notice that only pressure and moles are changing. Cross out volume and temperature from Ladon’s Law to reveal the equation you need to solve the problem:

The equation required is: (name, if any, is unknown to me!)

Assign the pressures and moles to the variables:

 P1 = 3.55 atm n1 = 15.0 mol

 P2 = 9.45 atm n2 = ?

11. (continued)

Rearrange and solve for n2 by cross-multiplying the equation and then dividing through by P1:

Cross multiplication yields: P1n2 = P2n1

Dividing through by P1 gives the equation needed to solve the problem:

Notice how the pressure units of atmospheres cancel out to leave mole units.

**12. What is the density of 55.2 grams of argon gas at a pressure of 3.95 atmospheres and a temperature of 35°C?**

The density of a gas is usually expressed in units of grams/liter, (g/L), which is a mass/volume. From the information in the problem, the mass is known, but the volume must be calculated from the ideal gas law. Let’s gather the information we need ( moles, pressure and temperature). Since we are using the ideal gas law, the pressure must be in atmospheres and the temperature must be in Kelvin. The moles of argon will be found using the atomic weight of argon, which is 39.95 g/mol.

 P = 3.95 atm (already in atmospheres, no conversion needed)

 V = ? L

 R = 0.08206 L∙atm/mol∙K

 T = 35°C + 273 = 308 K

Rearrange PV = nRT and solve for volume by dividing through be P:

Notice the pressure, mol, and temperature units cancel to leave volume units of liters. I’ll show you this as a complex fraction:

Use the mass and volume of the argon to find the density:

**13. A mixture of gas have the following partial pressures: 233 torr N2, 353 torr of He and 832 torr of O2. (a) What is the total pressure of the mixture of gases? (b) What are the mole fractions of each of these gases in the mixture?**

(a) What is needed here is Dalton’s Law of Partial Pressures. Dalton’s Law states that the total pressure is equal to sum of the partial pressures of all the gases in a mixture.

(b) A fraction is a “part” over a “whole”. The mole fraction, X, represents the moles of an individual gas divided by the total moles of all the gases in the mixture. Since moles and pressure are directly proportional, we can use the partial pressures and total pressure in place of moles:

Notice mole fraction does not have any units; it is a fraction.

**14. A mixture of gases contains 3.44 mol of He, 1.39 mol of Ne and 7.33 mol of N2. (a) What are the mole fractions of each of the gases in the mixture? (b) What are the partial pressures of each of these gases if the total pressure is 5.87 atmospheres?**

(a) Here we will use the moles of each gas to find the mole fraction of each gas. The total moles of gas is the sum of the individual moles:

The mole fraction, X, represents the moles of an individual gas divided by the total moles of all the gases in the mixture.

Notice once again, that mole fractions do not have units.

14. (continued)

(b) The partial pressures of each gas can be calculated by multiplying the total pressure by the mole fraction of each gas:

**15. An antacid tablet is analyzed for %CaCO3 by measuring the pressure, volume and temperature of the CO2 generated by adding acid to the sample according to the reaction:**

 **CaCO3(s) + 2HCl(aq) → CaCl2(aq) + CO2(g) + H2O(l)**

**A 0.200 gram tablet of antacid produces 34.55 mL of gas at 26.80°C. The gas was collected over water. Water at 26.8°C has a vapor pressure of 26.43 torr. The atmospheric pressure at the time of the analysis was 758.33 torr. What is the %CaCO3 in the antacid sample?**

This is fundamentally a stoichiometry problem. Most stoichiometry problems follow the strategy:

Quantity A → mol A → mol B → Quantity B

We will use this strategy to solve this problem. Our first task is to find the moles of carbon dioxide. Previously you have seen grams being converted to moles using a molar mass or a molarity and volume being used to find moles. Here the moles of gas is found by using the ideal gas law, PV = nRT.

Collect and convert the pressure, volume an temperature to the proper units. The pressure of the carbon dioxide alone must be found using Dalton’s Law:

Next, convert the pressure of carbon dioxide into atmospheres:

Convert the volume of carbon dioxide from milliliters to liters:

Convert the temperature in Celsius to Kelvin:

15. (continued)

We are now ready to find the moles of carbon dioxide using the ideal gas law. Rearrange PV = nRT to solve for moles by dividing through by RT:

We have just completed the first part of the stoichiometry strategy (Quantity A to mol A.) Let’s complete the remainder of the strategy. We will convert moles of carbon dioxide to moles of calcium carbonate to grams of calcium carbonate and finally to percent calcium carbonate:

Here is the problem summarized according the stoichiometry strategy:

 Quantity A → mol A → mol B → Quantity B

P, V. T of CO2 → mol CO2 → mol CaCO3 → g CaCO3 to %CaCO3

In the last step, we divided the grams of CaCO3 by the grams of antacid table and multiplied by 100 to obtain a percent. A percent is the quantity of a “part” divided by a “whole” multiplied by 100. See how the units in the above expression relate this!

**16. What is the density of sulfur dioxide gas at 974 torr and a temperature of 55.7 °C?**

From the ideal gas law we can derive an expression that relates, molar mass, pressure, density and temperature. Start with PV = nRT.

The moles, n, can be substituted with mass in grams (g) divided by molar mass, :

Multiply through by molar mass:

Divide by volume, V:

Since density, D, is mass divided by volume (g/V), substitute D for g/V:

An easy way to remember this equation is the mnemonic “Mud Pies are made from DiRT”.

16 (continued)

Using this equation, collect your terms and convert them to the appropriate unit:

 = 64.07 g/mol

 D = ? g/L

 R = 0.08206 L∙atm/mol∙K

 T = 55.7 °C + 273.15 = 328.8 K

Rearrange and solve for density:

**17. What pressure is needed to have neon gas at 56.9 °C have a density of 0.555 g/L?**

We will use to solve this problem. Collect terms and convert the temperature to the appropriate unit:

 P = ? atm

 D = 0.555 g/L

 R = 0.08206 L∙atm/mol∙K

 T = 56.9 °C + 273.15 = 330.0 K

Rearrange to solve for P by dividing through by the molar mass,

**18. An unknown gas has a density of 1.39 g/L at a temperature of 56.8°C and a pressure of 1.35 atmospheres. What is the molar mass of the gas?**

We will use to solve this problem. Collect terms and convert the temperature to the appropriate unit:

 R = 0.08206 L∙atm/mol∙K

 P = 1.35 atm T = 56.8 °C + 273.15 = 330.0 K

 D = 1.39 g/mol

18. (continued)

Rearrange to solve for the molar mass, by dividing through by pressure, P:

**19. An unknown gas with a mass of 2.56 g is in a 5.00 L container with a temperature of 345 K and a pressure of 0.0445 atmospheres. What is the molar mass of the gas?**

First find the density of the gas:

Rearrange to solve for the molar mass, by dividing through by pressure, P:

**20. A gas has a density of 1.39 g/L a56.8°C and a pressure of 1.35 atmospheres. What density will the gas have if the pressure is changed to 5.33 atmospheres and temperature remains constant?**

We can manipulate in the same way as PV = nRT. Only density and pressure are changing, so we can set up two expressions for each set of conditions:

 and

Rearrange both equations so that the variables that are changing are on one side of the expression and the variables that are constant are on the other side of the expression. To do this, divide both sides by the molar mass, , and divide through both sides of the expression by the density term to get:

Since both pressure/density expressions now equal to the same constant, they are equal to each other:

Cross-multiply and solve for D2 by dividing through by P1:

 becomes

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