

CH301 Worksheet 8—Gases (Answer Key)

1. What do we assume about ideal gases? What is the ideal gas law? Give the units for each variable.

Ideal gases are infinitely small, hard spheres that do not interact with each other. They are essentially "blind" to other gas molecules and will bounce off of each other just as they would bounce off of a wall.

The ideal gas law is $PV = nRT$. Common units for pressure, volume, n , and temperature are atmospheres, liters, moles, and Kelvin, respectively. R is a constant, which has a value of $0.08206 \text{ (L}\cdot\text{atm)/(K}\cdot\text{mol)}$.

2. If you know the number of moles of an ideal gas, what is the minimum number of variables that you need to know in order to fully determine the system?

If you know the number of moles of an ideal gas, you only need to know 2 more variables to recreate the system. Any two out of pressure, volume, and temperature is enough to define the final unknown term.

3. Assuming a constant molar quantity of gas, how could you produce the following effects?

a. decrease pressure

You could decrease the temperature or increase the volume.

b. decrease volume

You could decrease the temperature or increase the external pressure.

c. increase pressure

You could increase the temperature or decrease the volume.

d. increase volume

You could increase the temperature or decrease the external pressure.

4. You are scuba diving in a large fish tank. While you are at the bottom of the tank, you release a balloon full of air and watch it as it rises to the surface. What do you notice about the volume of the balloon?

The pressure at the bottom of the tank is greater than the pressure at the top of the tank, so as it rises, you notice that its volume increases.

5. What is the temperature of .75 moles of argon in a 18 L container with a pressure of 790 Torr?

$n = .75 \text{ mole}, V = 18 \text{ L}, \text{ and } P = 790 \text{ Torr } (1 \text{ atm}/760 \text{ Torr}) = 1.04 \text{ atm}$

$T = PV / (nR) = (1.04 \text{ atm} * 18 \text{ L}) / (.75 \text{ mol} * .082 \text{ Latm/molK}) = 304.4 \text{ K}$

6. Rank the following in order of increasing density:

1 mole of CH₄ at .1 atm and 273 K

2 moles O₂ at 1 atm and 300 K

3 mol H₂ 3 atm and 290 K

CH₄ < H₂ < O₂

CH₄:

$\rho = PMW / RT$

$= (.1 \text{ atm})(16\text{g/mol}) / (0.082 \text{ Latm/Kmol})(273\text{K}) = 0.07147 \text{ g/L}$

O₂

$\rho = PMW / RT$

$= (1 \text{ atm})(32\text{g/mol}) / (0.082 \text{ Latm/Kmol})(300\text{K}) = 1.3 \text{ g/L}$

H₂

$\rho = PMW / RT$

$= (3 \text{ atm})(2\text{g/mol}) / (0.082 \text{ Latm/Kmol})(290\text{K}) = 0.2523 \text{ g/L}$

7. You are a little boy or girl and stole your parents' hot air balloon. You are also a thirsty kid and brought some cans of soda (each can holds 354 mL) with you. As you get higher in altitude you notice that your soda cans start to expand and then eventually explode. So now you are thirsty and desperate to figure out what happened. You find out that the pressure at sea level is 1 atm and where the cans exploded is .7 atm. You also note that at sea level the temperature is 30 °C and 25 °C where the cans exploded. Why did the cans explode?

Without any calculations you should be able to do this problem. Note that the pressure dropped .3 atm. Pressure and volume are inversely related so volume must be increasing. Cans can only hold so much volume, hence why they exploded.

Now to prove it with some math:

$PV = nRT$

at sea level:

$$n = PV/RT$$

$$n = (1 \text{ atm})(.354\text{L}) / (0.082 \text{ Latm/Kmol})(273 + 30) = 0.014 \text{ moles}$$

in the air:

$$V = nRT/P = (.014 \text{ moles})(0.082 \text{ Latm/Kmol})(273 + 25) / .7\text{atm} = .4887 \text{ L}$$

So that means the liquid expanded from 354 mL to 489 mL.

8. An adult's lungs can hold about 6L. How many grams of air can an adult hold at a pressure of 102 kPa? Normal body temperature is 37 °C and air is about 20% oxygen and 80% nitrogen. (101,325 Pa = 1 atm)

Air is 20% O₂ and 80% N₂, oxygen and nitrogen are diatomic molecules

Oxygen: 32 g/mol

Nitrogen: 28 g/mol

Oxygen contributes 6.4 g/mol of air and nitrogen contributes 22.4 g/mol of air. So, air has a total of 28.8 g/mol.

$$(1 \text{ atm} / 101,325 \text{ Pa})(102000 \text{ Pa}) = 1.007 \text{ atm}$$

$$37 \text{ °C} + 273 = 310 \text{ K}$$

$$PV = nRT$$

$$n = PV/RT$$

$$n = (1.007 \text{ atm})(6\text{L}) / (0.082 \text{ Latm/Kmol})(310 \text{ K})$$

$$n = 0.238 \text{ moles}$$

$$(0.238 \text{ moles})(28.8 \text{ grams} / \text{mole}) = 6.85 \text{ grams of air}$$

9. Is it possible for 1 mole of air in an adult's lungs to be at STP? Explain and prove by use the ideal gas law.

No it's not possible, you would die. 1 mole of an ideal gas at STP has a volume of 22.4 L. This would be way too much for your lungs, which hold about 6L.

$$PV = nRT$$

At STP: 1 atm and 273K

$$V = nRT/P$$

$$V = (1 \text{ mole})(0.082 \text{ Latm/Kmol})(273\text{K}) / 1 \text{ atm}$$

$$V = 22.386 \text{ L}$$

10. A gas exerts a pressure of 1.12 atm in a 4 L container at 19C. You know the density of the gas is 1.5 g/L. What is the molecule?

$$PV = nRT$$

$$(1.12 \text{ atm})(4\text{L}) = n (0.082 \text{ Latm/Kmol})(19+273) = .187$$

$$(0.187\text{g}/4\text{L})(x \text{ g/mol}) = 1.5 \text{ g/L}$$

$$x = 32\text{g/mol}$$

O₂

11. What assumptions do we make when using the ideal gas law? Which of these are pretty good and which are pretty bad?

Gas molecules are infinitely small (i.e. euclidean points) - this is a very bad assumption, since all gases, even the smallest ones, have some volume.

The energy contained in a gaseous system is determined only by the absolute temperature - this is a very good assumption and is true over a wide range of temperatures.

All collisions (gas-gas or gas-container) are fully elastic - this is a very bad assumption.

There are no attractive or repulsive forces between molecules - this is also a very bad assumption, if it were true everything would be in the gas phase at all temperatures.

There are many more, but these are the ones we care about most.

12. What are the the two "fudge factors" (aka coefficients) in the van der Waals equation? Which terms in the ideal gas law do they correct/modify?

These "fudge factors" are "a" which corrects the pressure term of the ideal gas law and "b" which corrects the volume term in the ideal gas law.

13. Rank the following gases from most to least ideal in terms of the van der Waals coefficient b: CO₂, SF₆, O₂, H₂, He, CH₄, Rn,

$$H_2 > He > CH_4 > O_2 > CO_2 > SF_6 > Rn$$

Since coefficient b corrects for the volume occupied by the gas, larger gases will be less ideal.

14. At a given temperature, what will be the ratio of the rate of effusion of ozone to rate of effusion of molecular oxygen?

$$(m_1/m_2) = (v_2/v_1)^2$$

$$(m_{\text{oxygen}}/m_{\text{ozone}})^{1/2} = (v_{\text{ozone}}/v_{\text{oxygen}})$$

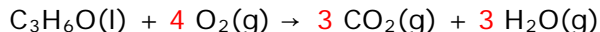
$$(32/48)^{1/2} = (v_{\text{ozone}}/v_{\text{oxygen}})$$

$$0.82:1 = v_{\text{ozone}}:v_{\text{oxygen}}$$

15. What is meant by STP? What are its values? Do you need to memorize these values?

STP is standard temperature and pressure; for the purposes of this course, standard temperature is 273 Kelvin (K), and standard pressure is 1 atmosphere (atm). Yes.

16. Balance the reaction below. If it goes to completion, what total volume will it occupy at STP? Will the volume have increased or decreased? By how much? (Assume ideality.)



3 moles of carbon dioxide and 3 moles of water vapor will give 6 total moles once the reaction goes to completion. A mole of ideal gas occupies ~22.4 L at STP. The total volume will increase from 89.6 L by 44.8 L to 134.4 L

17. If a gas molecule is moving at 800 miles per hour at a given temperature, by what factor would we need to increase the temperature in order to double the velocity of the gas molecule?

$$(T_2/T_1)^{1/2} = (v_2/v_1)$$

$$(xT_1/T_1) = (2/1)^2$$

One would need to increase the temperature by a factor of 4.

18. Why are diffusion and effusion so much slower than the actual velocity of a gas, (e.g. a gas molecule moving at 1000 km/hr diffuses at a tiny fraction of that rate)?

The velocity of gas describes the rate at which it moves, which, as a consequence of collisions, is rarely in a straight line. Diffusion and effusion describe net directional motion, which must be slower since so much of gas molecule's very fast motion is wasted moving backward.

19. An unidentified gas has a velocity of 753 m·s⁻¹ at STP. What is the identity of this gas? (Yes, this problem contains enough information to answer the question. Yes, it is hard.)

$$kT = mv^2$$

$$1.38\text{E}10^{-23} \cdot 273 = m \cdot 753^2$$

$$m = 6.64\text{E}10^{-27} \text{ kg}\cdot\text{molecule}^{-1}$$

$$6.64\text{E}10^{-27} \text{ kg}\cdot\text{molecule}^{-1} (1000 \text{ g/kg})(6.022\text{E}10^{23} \text{ molecules/mole}) = 4.001 \text{ g}\cdot\text{mol}^{-1}$$

The unknown gas is therefore helium.

20. Rank the following gases from least to most ideal in terms of the van der Waals coefficient a: N₂, H₂, HCl, HF, NH₃

$$\text{HF} < \text{NH}_3 < \text{HCl} < \text{N}_2 < \text{H}_2$$

Since coefficient a corrects for the pressure exerted by the gas, species with greater IMF will be less ideal.