Watch Bozeman Videos & other videos on my website for additional help:
Big Idea 2:  
• Gases

10.1 Characteristics of Gases
Read p. 398-401. Answer the Study Guide questions
1. Earth’s atmosphere is made up of many gases. Which 2 gases make up the most of the atmosphere? What are their percent’s?
2. What is the difference between a gas and vapor?
3. Determine what happens when pressure is applied to gas?
4. Describe how different gases behave similar to each other?

<table>
<thead>
<tr>
<th>Shape</th>
<th>Volume</th>
<th>Compressible</th>
<th>Flow/Speed</th>
<th>Type of mixtures formed</th>
<th>Density units</th>
</tr>
</thead>
<tbody>
<tr>
<td>Solids</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
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<tr>
<td>Liquids</td>
<td></td>
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<tr>
<td>Gas</td>
<td></td>
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</tr>
</tbody>
</table>
Gases are made up of particles that have (relatively) large amounts of energy.
Gases highly compressible – when pressure is applied to a gas, its volume decreases.
A gas has no definite shape or volume and will expand to fill as much space as possible.
As a result of the large amount of empty space in a volume of gas, gases are easily compressed.

10.2 Pressure
Read p. 401-404. Answer the Study Guide questions

1. Define Pressure
2. What is the SI Unit for Pressure
3. Describe in own terms how pressure is measured in chemistry and what is filled inside of it?
4. Draw Figure 10.3 (p.403)
5. Determine why gas particles override the gravitational forces?
6. Identify what a Pascal, Newton, and a bar have and don’t have in common?
7. Determine who identified that the atmosphere has mass and how he identified it?
8. Identify the different types of standard atmospheric pressure numbers and units? (The 5 “magic numbers” of pressure p.402)

Pressure is exerted when the gas particles collide with the walls of any container it is held in.

Atmospheric Pressure and the Barometer
• The SI unit of pressure is the pascal (Pa).

• Atmospheric pressure is measured with a barometer.
  • If a tube is completely filled with mercury and then inverted into a container of mercury open to the atmosphere, the mercury will rise 760 mm up the tube.

Important non SI units used to express gas pressure include:
  atmospheres (atm)
  torr (torr)
  millimeters of mercury (mm Hg)
  Pascal (Pa)
  kilopascal (kPa)

1 atm = 760. mm Hg = 760. torr = 101, 325 Pa = 101.325 kPa
Examples
Perform the following conversions
1. 657 mmHg to torr

2. 830 torr to atmospheres

3. 7.8 kPa to atmospheres

4. 1200 Pa to mmHg

Barometer with Hg Pic (We will draw 3 different pics here)

Example
1) The height of the mercury in the open ended arm is 136.4mm, and the height in the arm in contact with the gas in the flask is 103.8mm. What is the pressure of the gas in the flask (a) in atmospheres, (b) in torr. (Draw and Label the barometer)
10.3 The Gas Laws

Read p. 404-407. Answer the Study Guide questions

1. Identify the four variables involved to define the physical condition of a gas?
2. Does atmospheric pressure increase or decrease as altitude increases? (Neglect changes in temperature) Explain. (See figure 10.4)
3. Determine the effect pressure has on gas as it increases or decreases?
4. What is Boyle’s Law?
5. What is the equation for Boyle’s Law?
6. What variable must stay constant to apply Boyle’s Law?
7. According to Boyle’s Law, ______________ is ______________ proportional to ______________.
8. Sketch a graph of the relationship between V and P for Boyle’s Law? See figure 10.6
9. What is Charles’s Law?
10. What is the effect of temperature (warm/cold) on volume? See Figure 10.7
11. What is the equation for Charles’s Law?
12. What units does Temperature have to be in?
13. What variable must stay constant to apply Charles’s Law?
14. According to Charles’s Law, ______________ is ______________ proportional to ______________.
15. Sketch a graph of the relationship between V and T for Charles’s Law? See figure 10.7
16. If temperature increases, gas particles gain ______________ ______________
17. What is Absolute Temp?
18. How to you convert Celsius to Kelvin?
19. What is Gay-Lussac’s Law?
20. What is the equation for Gay-Lussac’s Law?
21. What units does Temperature have to be in?
22. What variable must stay constant to apply Gay-Lussac’s Law?
23. According to Gay-Lussac’s Law, ______________ is ______________ proportional to ______________.
24. Sketch a graph of the relationship between P and T for Gay-Lussac’s Law?
25. If temperature increases, gas will have more ______________ and ______________ more with each other and the walls of the container.
26. What is Avogadro’s Law?
27. What is the equation for Avogadro’s Law?
28. What variable must stay constant to apply Avogadro’s Law?
29. According to Avogadro, ______________ is ______________ proportional to ______________.
The equations that express the relationships among $T$ (temperature), $P$ (pressure), $V$ (volume), and $n$ (number of moles of gas) are known as the gas laws.

The Pressure-Volume Relationship: Boyle’s Law

Boyle’s Law states that at constant temperature and moles, pressure is inversely proportional to volume. This means that as the pressure increases the volume decreases and vice versa.

This makes sense. If the volume is increased, the gas particles collide with the walls of the container less often and the pressure is reduced.

\[ P_1 V_1 \rightarrow P_2 V_2 \]

$P_1$ and $V_1$ are the original conditions. $P_2$ and $V_2$ are the new conditions.

A plot of $P$ versus $V$ is a hyperbola.

- The working of the lungs illustrates that:
  - as we breathe in, the diaphragm moves down, and the ribs expand; therefore, the volume of the lungs increases.
  - according to Boyle’s law, when the volume of the lungs increases, the pressure decreases; therefore, the pressure inside the lungs is less than the atmospheric pressure.
  - atmospheric pressure forces air into the lungs until the pressure once again equals atmospheric pressure.
  - as we breathe out, the diaphragm moves up and the ribs contract; therefore, the volume of the lungs decreases.
  - By Boyle’s law, the pressure increases and air is forced out.

Examples

1. If a 1.23 L sample of a gas at 53.0 torr is put under pressure up to a value of 240. torr at a constant temperature, what is the new volume?
The Temperature-Volume Relationship: Charles’s Law
Charles’ Law states that, at constant pressure and moles, volume is directly proportional to temperature.

This means the volume of a gas increases with increasing temperature and vice versa.

If the temperature is increased, the gas particles gain kinetic energy, move around more and occupy more space.

\[
V_1 \, T_2 \rightarrow V_2 \, T_1
\]

*** \(T\) needs to be in KELVIN.

Absolute zero, 0 K = –273 °C.

\(V_1\) and \(T_1\) are the original conditions. \(V_2\) and \(T_2\) are the new conditions.

A plot of \(V\) versus \(T\) is a straight line.

Examples
1. An 11.0 L sample of a gas is collected at 276 K and then cooled by 14 K. The pressure is held constant at 1.20 atm. Calculate the new volume of the gas.
**Pressure and Temperature relationships: Gay-Lussac’s Law**

Gay-Lussac’s Law states that, at constant volume and moles, pressure is directly proportional to temperature.

This means that temperature increases with increasing pressure and vice versa.

If the temperature of a gas is raised then the particles will have more energy and collisions with the walls of the container will be more forceful and the pressure will increase.

\[ P_1 \ T_2 \rightarrow P_2 \ T_1 \]

*** T needs to be in KELVIN.***

**Absolute zero, 0 K = –273 °C.**

*P1 and T1 are the original condition. P2 and T2 are the new conditions.*

A plot of P versus T is a straight line.

**Examples**

1. A gas at 25°C in a closed container has its pressured raised from 150. atm to 160. atm. What is the final temperature of the gas?
The Quantity-Volume Relationship: Avogadro’s Law

Avogadro’s Law states that, at constant temperature and pressure, volume is directly proportional to the number of moles of gas present.

This means the volume of a gas increases with increasing number of moles and vice versa.

As more moles of a gas are placed into a container if conditions of temperature and pressure are to remain the same, the gas must occupy a larger volume.

\[ V_1 \cdot n_1 \rightarrow V_2 \cdot n_2 \]

V₁ and n₁ are the original condition. V₂ and n₂ are the new conditions.

Can also use Stoichiometry!!! (See 10.4 why we can)

According to Figure 10.9 - Draw the 3 gases for each information below.

How many moles of gas are in each vessel? THINK

<table>
<thead>
<tr>
<th>Volume:</th>
<th>22.4L</th>
<th>22.4L</th>
<th>22.4L</th>
</tr>
</thead>
<tbody>
<tr>
<td>Pressure:</td>
<td>1atm</td>
<td>1atm</td>
<td>1atm</td>
</tr>
<tr>
<td>Temperature (Celsius):</td>
<td>0</td>
<td>0</td>
<td>0</td>
</tr>
<tr>
<td>Mass of Gas:</td>
<td>4.00g</td>
<td>28.0g</td>
<td>16.0g</td>
</tr>
<tr>
<td># of gas molecules:</td>
<td>(6.02 \times 10^{23})</td>
<td>(6.02 \times 10^{23})</td>
<td>(6.02 \times 10^{23})</td>
</tr>
</tbody>
</table>

Examples

1. A 13.1 L sample of 0.502 moles of \(O_2\) is held under conditions of 1.00 atm and 25.0 °C. If all of the \(O_2\) is then converted to Ozone (\(O_3\)) what will be the volume of ozone?
10.4 Combined Gas Law & The Ideal-Gas Equation

Read p. 408-412. Answer the Study Guide questions

1. What is Combined Gas Law?
2. What is the equation?
3. What units do each variable have to be?
4. What variable must stay constant to apply Combined Gas Law?
5. What is Ideal Gas Law?
6. What is the equation?
7. What units do each variable have to be?
8. What is R called and what are the different values/units for it? See Table 10.2
9. What is STP?
10. What is molar volume and what temperature does it occur at?
11. For what state of matter can you only use molar volume for?
12. Identify the numbers of volume and temperature at STP?
13. Compare and contrast an ideal gas to a real gas and how you can’t use the ideal gas law on a real gas?

Relating Boyle's, Charles, Gay-Lussac, & Avogadro's Laws: COMBINED Gas Laws

- You can combine the four gas laws into one equation:

\[
\frac{P_1V_1}{n_1T_1} = \frac{P_2V_2}{n_2T_2}
\]

**Ideal gas equation:** An ideal gas is a hypothetical gas whose P, V, and T behavior is completely described by the ideal-gas equation.

*** Have to have certain units to plug into equation***

\[
PV = nRT
\]

- P = pressure (atm)
- V = volume (L)
- n = number of moles (mol)
- T = temperature (K)
- R = gas constant = 0.08206 L · atm/mol · K

- Other numerical values of R in various units are given in Table 10.2.
Rearrange Ideal Gas Equation to solve for each variable:

\[ P = \quad V = \quad n = \quad T = \]

**Example**

1. Assuming ideal behavior, how many moles of Helium gas are in a sample that has a volume of 8.12 L at a temperature of 0.00 °C and a pressure of 1.20 atm?

2. The atmospheric pressure of 4.4X10^{-2} mol of CH_4 is 0.599 atm at 331 °C. Determine the volume of CH_4 under these circumstances.

**** Stoichiometry:

22.4 L of any gas at 0 °C (273K) & 1 atm contains \(6.02 \times 10^{23}\) gas molecules = 1 mole.

YOU CAN ONLY USE 22.4L FOR GASES!!!! (Works with Avogadro’s Law)

STP (standard temperature and pressure) = 0°C or 273 K at 1 atm.

- The **molar volume** of 1 mol of an ideal gas at STP is 22.41 L.

**Example:**

1. Calculate the mass of ammonium chloride required to produce 11.6 L of ammonia gas in the reaction below.

\[ 2NH_4Cl(s) + Ca(OH)_2(s) \rightarrow 2NH_3(g) + CaCl_2(s) + 2H_2O(g) \]
10.5 Further Applications of the Ideal-Gas Equation

Read p. 412-415. Answer the Study Guide questions

1. Identify what the ideal Gas equation can be used to identify?
2. Why is there not a equation for real gases?
3. Describe how you can calculate density by using the ideal gas equation (Unless you want to memorize, derive it)?
4. Describe how you can calculate molar mass by using the ideal gas equation (Unless you want to memorize, derive it)?
5. Compare and contrast how molar mass changes the denseness of gases?
6. Determine the importance of understanding the properties of gases?

Gas Densities and Molar Mass

- Density has units of mass over volume (Gases use density units of grams/liters).
- Rearranging the ideal-gas equation with \( M \) as molar mass (g/mol) we get:

\[
\frac{n}{V} = \frac{P}{RT}
\]

\[
\frac{nM}{V} = \frac{PM}{RT}
\]

\[
d = \frac{PM}{RT}
\]

- The molar mass of a gas can be determined as follows:

\[M = \frac{dRT}{P}\]

Examples

1. A sample of aluminum chloride weighing 0.100 g was vaporized at 350. \(^\circ\)C and 1.00 atm pressure to produce 19.2 mL of vapor. Calculate a value for the Molar Mass of aluminum chloride.
2. The density of a certain molecule is 3.164 (g/L) at a constant pressure of 1.0 atm in 0°C. What is the molecule’s molar mass, and identify the molecule?

10.6 Gas Mixtures and Partial Pressures
Read p. 415-418. Answer the Study Guide questions

1. How do we deal with gases composed of a mixture of two or more different substances?
2. What is Dalton’s Law?
3. What is the equation?
4. What units do each variable have to be?
5. What variable must stay constant?
6. Describe how Dalton’s law is used to identify the total pressure of a mixture?
7. Compare and contrast the partial pressure’s to mole fractions?
8. What is mole fractions equation?

- **Dalton observed:**
  - The total pressure of a mixture of gases equals the sum of the pressures that each would exert if present alone as long as volume, temperature, and moles are constant!!!
  - **Partial pressure** is the pressure exerted by a particular component of a gas mixture.
  - **Dalton’s law of partial pressures:** In a gas mixture the total pressure is given by the sum of partial pressures of each component:
    \[
    P_t = P_1 + P_2 + P_3 + \ldots
    \]

Since gas molecules are so far apart, we can assume that they behave independently.

- Each gas obeys the ideal gas equation:
  \[
  P_i = \left( n_i + n_2 + n_3 + \ldots \right) \frac{RT}{V} = n_i \frac{RT}{V}
  \]
Partial Pressures and Mole Fractions
If you wanted to solve for the pressure of only 1 gas in a mixture of other gases....

\[ P_1 = X_1 P_t \]

\( P_1 \) = pressure for a certain gas  
\( X_1 \) = mole fraction \( \left( \frac{n_1}{n_t} \right) \) 
\( P_t \) = total pressure of all gases in container

If you substitute mole fraction \( \left( \frac{n_1}{n_t} \right) \) into the equation for partial pressure, you get the following:

\[ P_1 = \left( \frac{n_1}{n_t} \right) P_t \]

Mole percent = mole fraction x 100
So if you have 80 mol percent of Ar, then its mole fraction would be .80

Examples

a. A synthetic atmosphere is created by blending 2 mol percent CO₂, 20 mol percent O₂ and 78 mol percent N₂. If the total pressure is 750 torr, calculate the partial pressure of the oxygen component.

b. If 25L of this atmosphere, at 37° C, have to be produced, how many moles of O₂ are needed?
10.7 Kinetic-Molecular Theory – VERY IMPORTANT!!!
Read p. 418-421. Answer the Study Guide questions

1. The ____________ ___________ equation determines how gases behave but not why they behave as they do.
2. Why does a gas expand when heated at constant pressure?
3. Why does its pressure increase when the gas is composed at constant temperature?
4. Summarize what the Kinetic Molecular Theory (KMT) model is for gases.
5. THE KMT explains both ____________ and ______________ at the molecular level.
6. How does pressure increase inside a container?
7. The magnitude of the pressure is determined by how?
8. If 2 gases are at the same temperature, their molecules have the same
   ____________ ____________ ____________ ____________
9. Molecular motion increases with ____________ ____________ ____________
10. Describe how collisions of molecules may increase or cause the molecules to have constant kinetic energy?
11. The molecules in a sample of gas have an ____________ ____________ ____________ ____________ ____________, but the individual molecules are moving ____________ ____________ ____________ ____________ ____________.
12. Read p.419 really well and Figure 10.13a. Draw this figure. Understand the red vs. blue line. Which has the higher average kinetic energy?
13. What is root-mean square (rms) speed?
14. Because mass does not change with temperature, the average kinetic energy ____________ ____________ as the temperature ____________. This means that the ____________ increases as Temperature ____________ (This is huge that you understand this)
15. Determine the effect of a volume increase at constant temperature? (3 things happen)
16. Describe the effect of a temperature at constant volume? (5 things happen)
17. According to Sample Exercise 10.12 p.420, A sample of O2 gas initially at STP is compressed to a smaller volume at constant temperature. What effect does this change have on
   (a) the average kinetic energy of the molecules
   (b) their average speed
   (c) the number of collisions the make with the container walls per unit time
   (d) the number of collisions they make with a unit area of container wall per unit time
   (e) the pressure?
• The **Kinetic-Molecular theory (KMT)** was developed to *explain* gas behavior.
  • It is a **theory of moving molecules**.

**Summary: (KNOW THIS WELL)**
1. Gases consist of a large number of molecules in constant random motion.
2. The combined volume of all the molecules is negligible compared with the volume of the container.
3. Intermolecular forces (forces between gas molecules) are negligible.
4. Energy can be transferred between molecules during collisions, but the average kinetic energy is constant at **constant temperature**.
5. The collisions are perfectly elastic.
6. The average kinetic energy of the gas molecules is proportional to the absolute temperature.

• Kinetic molecular theory gives us an *understanding* of pressure and temperature on the molecular level.
  • The pressure of a gas results from the collisions with the walls of the container.
  • The magnitude of the pressure is determined by how often and how hard the molecules strike.
    - Some molecules will have less kinetic energy (KE) or more kinetic energy than the average
      • As the temperature increases, the average kinetic energy of the gas molecules increases.
      • As kinetic energy increases, the velocity of the gas molecules increases.

**Application to the Gas-Laws**
• The **effect of an increase in volume (at constant temperature)** is as follows:
  • As volume increases at constant temperature, the average kinetic energy of the gas remains constant.
  • Therefore, $u$ is constant.
  • The gas molecules have to travel further to hit the walls of the container.
  • Therefore, pressure decreases. (Boyle's Law)

• The **effect of an increase in temperature (at constant volume)** is as follows:
  • If temperature increases at constant volume, the average kinetic energy of the gas molecules increases.
  • There are more collisions with the container walls.
  • Therefore, $u$ increases.
  • The change in momentum in each collision increases (molecules strike harder).
  • Therefore, pressure increases. (Gay-Lussac Law)
10.8 Molecular Effusion and Diffusion

Read p. 421-426. Answer the Study Guide questions

1. According to KMT, for 2 gases at the same T a gas composed of ______________ particles, such as He, will have the same ______________ ______________ as one composed of ______________ particles, such as Xe.

2. According to figure 10.14 (read p.422), which of these gases has the largest molar mass? List them from smallest to largest.

3. What is root-mean square (rms) speed?
4. What is the symbol for it?
5. What is the equation?
6. What units does each variable have to be?
7. What unit does R have to be?....(Be careful)
8. Determine the differences in the two terms of effusion and diffusion?
9. Determine who identified the effusion rate of a gas?
10. What is Grahams Law?
11. What is the equation?
12. What units does each variable have to be?
13. How do you know which molar mass goes where in the equation?
14. Describe how a faster molecule may effuse faster than a slower molecule?
15. Identify what varies with the mean free path and give an analogy to describe it?
16. According to Figure 10.16 p. 424, because pressure and temperature are constant in this figure but volume changes, which other quantity in the ideal gas equation
must change? Compare two different balloons, one of Argon gas and one of Helium gas. Explain this process of effusion within the balloons.

• The average kinetic energy of a gas is related to its mass:

• Consider two gases at the same temperature: the lighter gas has a higher rms (speed) than the heavier gas.
  • Mathematically:

\[ u = \sqrt{\frac{3RT}{M}} \]

• \( R = 8.314 \text{ J/K} \cdot \text{mol} \)
  \( M = \text{molar mass in kg/mol} \)
  \( T = \text{Kelvin} \)

• Two consequences of the dependence of molecular speeds on mass are:
  • **Effusion** is the escape of gas molecules through a tiny hole into an evacuated space. (balloon)
  • **Diffusion** is the spread of one substance throughout a space or throughout a second substance.

**Examples**
1. Calculate the rms speed of the molecules in a sample of Helium and Nitrogen gas at 25°C.

   a. **Helium**

   b. **Nitrogen**
2. Consider 3 gases all at 298K: HCl, H₂, and O₂. List the gases in order of increasing average speed.

3. How is the rms speed of N₂ molecules in a gas sample changed by
   a. Increase in temperature
   b. Increase in volume
   c. Mixing with a sample of Ar at the same temperature.

Graham’s Law of Effusion
- The rate of effusion can be quantified.
- Consider two gases with molar masses, \( M_1 \) and \( M_2 \), and with effusion rates, \( r_1 \) and \( r_2 \), respectively.
  - The relative rate of effusion is given by Graham’s law:
    \[
    \frac{r_1}{r_2} = \sqrt{\frac{M_2}{M_1}}
    \]
  - Only those molecules which hit the small hole will escape through it.
  - Therefore, the higher the rms speed the more likely it is that a gas molecule will hit the hole.

Diffusion and Mean Free Path
- Diffusion is faster for light gas molecules.
- Diffusion is slowed by collisions of gas molecules with one another.
  - Consider someone opening a perfume bottle: It takes awhile to detect the odor, but the average speed of the molecules at 25°C is about 515 m/s (1150 mi/hr).

Examples
1. Calculate the ratio of effusion rate of N2 and O2 gases.
2. Calculate the ratio of effusion rate of $^{238}\text{UF}_6$ and $^{235}\text{UF}_6$ gases.

3. An unknown gas composed of homonuclear diatomic molecules effuses at a rate that is .355 times the rate at which Oxygen gas effuses at the same temperature. Calculate the molar mass of the unknown and identify it.

10.9 Real Gases: Deviations from Ideal Behavior – AP Free Response Question!!!!

Read p. 426-429. Answer the Study Guide questions

1. Describe why real gases don’t follow the ideal gas law?
2. Real Gases differ from ideal gases at ____________ ______________ and at ______________ ______________
3. Compare the volume of real and ideal gas molecules of a volume container?
4. Describe the effect of intermolecular forces on gas pressure?
5. Determine what comes into play when molecules are crowded together at high pressure?
6. Determine the equation scientist have to use to calculate for real gases?
7. Which gas deviates most from ideal behavior? Justify.
• From the ideal gas equation:

\[
\frac{PV}{RT} = n
\]

• For 1 mol of an ideal gas, \( PV/RT = 1 \) for all pressures.
  • The higher the pressure the more the deviation from ideal behavior.

• For 1 mol of an ideal gas, \( PV/RT = 1 \) for all temperatures.
  • As temperature increases, the gases behave more ideally.

THIS IS WHERE IDEAL GASES DIFFER FROM REAL GASES!!!!!!

Deviations from ideal behavior
At high pressures and low temperatures gas particles come close enough to one another to make the two postulates of the Kinetic Theory below, invalid.

(i) Gases are composed of tiny particles whose size is negligible compared to the average distance between them. – When the gas is pressurized into a small space the gas particles size becomes more significant compared to the total volume.

(ii) The forces of attraction or repulsion between two particles in a gas are very weak or negligible. – Low temperature means less energy, so the particles are attracted to one another more.

Under these conditions (high P and low T) gases are said to behave non-ideally or like “real” gases. This has two consequences.

(i) When gases are compressed to high pressures, the size of the gas particles is no longer negligible compared to the total space occupied by the gas (its total volume). Therefore, the observed total volume occupied by the gas under these real conditions is artificially large since the gas particles are now occupying a significant amount of that total volume.

(ii) The actual pressure of a gas is lower than one would expect when assuming there were no attractive forces between the particles. Because, in a real gas, the particles are attracted to one another, they collide with the walls with less force, and the observed pressure is less than it an ideal gas.