Unit 8 Packet: Gas Laws
Introduction to Gas Laws Notes:

- In chemistry, the relationships between gas physical properties are described as gas laws. Some of these properties are pressure, volume, and temperature. These laws show how a change in one of these properties affects the others. The gas laws in chemistry are: Boyle's Law, Charles' Law, the Combined Gas Law, Avogadro's Law, and the Ideal Gas Law.

- Gas Variables:
  - **Pressure** (P) – The force per unit area on a surface. Gas molecules exert force, and therefore pressure, on any surface, sides of container, with which they collide.
  - **Temperature** (T) – Temperature is measured in the Kelvin scale; it is an index of gas motion and not a measure of heat. Absolute zero, K = 0°C, all motion stops!
  - **Volume** (V) – Volume is the quantity of three-dimensional space occupied by a liquid, solid or gas. Common units used to express volume include liters and cubic meters.
  - **Moles** (n) – A mole is the amount of pure substance containing the same number of chemical units as there are atoms in exactly 12 grams of carbon-12 (i.e., 6.023 X 10^{23}).
### Equations Representing Relationships:

<table>
<thead>
<tr>
<th>Equation</th>
<th>Law</th>
</tr>
</thead>
<tbody>
<tr>
<td>$P_{\text{total}} = P_1 + P_2 + P_3 ...$</td>
<td>Dalton’s Law of Partial Pressure</td>
</tr>
<tr>
<td>$P_1V_1 = P_2V_2$</td>
<td>Boyle’s Law</td>
</tr>
<tr>
<td>$V_1 / T_1 = V_2 / T_2$</td>
<td>Charles’ Law</td>
</tr>
<tr>
<td>$V_1 / n_1 = V_2 / n_2$</td>
<td>Avogadro's Law</td>
</tr>
<tr>
<td>$P_1 / T_1 = P_2 / T_2$</td>
<td>Gay-Lussac’s Law</td>
</tr>
<tr>
<td>$P_1V_1 / T_1 = P_2V_2 / T_2$</td>
<td>Combined Gas Law</td>
</tr>
<tr>
<td>$PV = nRT$</td>
<td></td>
</tr>
<tr>
<td>$R = 8.3145 \text{ L kPa/mol K}$</td>
<td></td>
</tr>
<tr>
<td>$R = 0.08206 \text{ L atm/mol K}$</td>
<td></td>
</tr>
</tbody>
</table>

### Standard Temperature and Pressure (STP):

Standard temperature is defined as zero degrees Celsius ($0^\circ C$), which translates to 32 degrees Fahrenheit ($32^\circ F$) or 273.15 degrees Kelvin. Standard pressure is 101.325 kPa, 1 atmosphere (1 atm) or 760 millimeters in a mercurial barometer (760 mmHg) or 760 torrs.
Pressure and Kinetic Molecular Theory Notes

What is pressure?

\[ \text{Pressure} = \frac{\text{Force}}{\text{Area}} \]

What is force? You can think of it like weight or like a push. How much does your body push on the floor? How much do you have to push to move something across the floor?

Liquids and solids exert pressure.

Ex: Diving, your body is under more pressure when under water than when above it.

If you are trapped under a piece of furniture you feel the pressure of the furniture on you.

What does area have to do with it?

Could you sleep on a bed of nails? How about just one nail?

Carpenters…why are screws pointy at the end instead of flat?

The same force over a smaller area results in a higher pressure.

Gas pressure is similar but hard to visualize.

KINETIC MOLECULAR THEORY

1. Gases have a very small _attraction to each other_
2. Gases are constantly in _motion_ at high speeds in random but _straight_ line paths
3. Gases experience no _real_ attractions between _other gas molecules._
   a. All collisions between particles in a gas are perfectly _elastic_.
      i. No attractive or repulsive forces
      ii. No transfer of Kinetic Energy
      iii. The average Kinetic Energy is dependent only on _temperature_

4. The speed of a gas molecule is directly _proportional_ to its mass and speed.
   a. As the temperature increases, the speed of the gas molecule _increases._

\[ KE = \frac{1}{2} mV^2 \]
These assumptions are good for IDEAL gases and are used quite often with little error. However, when one needs to be exact that you must account for the fact that REAL gases

- Have volume
- Experience electrostatic attractions

For Introductory Chemistry, though, we will just deal with Ideal gases.

**Pressure** is the sum of all the forces of all the gas molecules colliding with a surface. Gas particles are in **constant random motion** exerting pressure as they collide with the walls of the container. Therefore, the **more** collisions, the **higher** the pressure.

Gases have certain properties that can be explained by the KMT.

1) **Low Density** they **have small volume** because the molecules are moving at a high rate of speed and are not held back by electrostatic attractions as well as being spread out, they have low density.

2) **Compressibility**, they can be **easy** to compress because the molecules have space between them, unlike liquids and solids where there is little space between the molecules.

3) **Expansion**, they will **spread** given the opportunity because the molecules are moving at a high rate of speed.

4) **Diffusion** they will spread out, again because the molecules are moving at a high rate of speed,
   a. Example: perfume diffusing through air
   b. Example: Liquids also diffuse: food coloring in water

5) **Effusion**, molecules moving at a high rate of speed will eventually “collide” with a hole and escape.
   a. Effusion is gas escaping through a hole,
b. Example: air escaping through a hole in your tire

1. The word *kinetic* comes from a Greek word that means “to move.” The **kinetic molecular theory** is based upon the assumption that particles of matter (atoms or molecules) are in constant motion.

2. Of the three states of matter, which one has the most kinetic energy? **gas**

3. Which state of matter has particles that are separated by the largest distance? **gas**

4. A **scientific theory** is an explanation of some type of natural phenomena. Theories are normally developed from careful study of the way the world behaves. Let’s look at how gases behave and see if the kinetic molecular theory makes sense.

5. Compared to liquids and solids, gases tend to have **lower** densities. This can be explained because the particles of gas are **far apart in constant motion**.

**Units of pressure**

- **Kilopascals** = SI units of pressure, named after scientist Pascal
- **Torr** = named after scientist Torricelli
- **mmHg** = millimeters of mercury, comes from old way of measuring pressure by the height of a column of mercury open to atmosphere
- **atm** = atmosphere, unit most often used in chemistry. One atmosphere is what we feel experience every day. It is the standard pressure on earth.
- **psi** = pounds per square inch. Most often used in real life (tire pressure)
Gas Laws Notes

Steps to Solve any Gas Law Problem:
- Step 1: Write everything you are given in the problem.
- Step 2: Which law do you want to use? (What remains constant?)
- Step 3: Do your units match? If not, convert. (Temperature must always be in Kelvin)
- Step 4: Plug in your values and solve.

- Boyle's Law
  \[ P_1 V_1 = P_2 V_2 \]
  - As the pressure decreases, the volume increases.
  - Indirectly proportional
  - Temperature remains constant

Example Problem: A balloon contains 30.0 L of helium gas at 1 atmosphere, and it rises to an altitude where the pressure is only 0.25 atm, assuming the temperature remains constant, what is the volume of the balloon at its new pressure?

\[
\begin{align*}
P_1 &= 1 \text{ atm} \\
V_1 &= 30.0 \text{ L} \\
P_2 &= 0.25 \text{ atm} \\
V_2 &= x
\end{align*}
\]

\[
(1 \text{ atm})(30.0 \text{ L}) = (0.25 \text{ atm})x
\]

\[
x = 120 \text{ L}
\]

- Charles' Law
  \[ \frac{V_1}{T_1} = \frac{V_2}{T_2} \]
  - Lower temperature leads to a lower volume
  - Higher temperature leads to a higher volume
  - Directly proportional
  - Temperature must be converted to Kelvin
  - Pressure remains constant
**Example Problem:** A balloon inflated in a room at 24°C has a volume of 4.00 L. The balloon is then heated to a temperature of 58°C what is the new volume if the pressure remains constant?

\[
\frac{V_1}{T_1} = \frac{V_2}{T_2}
\]

\[
\frac{4.00 \text{ L}}{297 \text{ K}} = \frac{x}{331 \text{ K}}
\]

\[
X = 4.46 \text{ L}
\]

- **Gay-Lussac's Law**

\[
\frac{P_1}{T_1} = \frac{P_2}{T_2}
\]

- Temperature always in Kelvin Scale
- Volume remains constant

Example Problem: Aerosol cans carry labels warning not to store them above a certain temperature. The gas in a used aerosol can is at a pressure of 1 atm at 25°C. If the can is thrown into a fire, what will the pressure be when the temperature reaches 1201°C?

\[
\frac{P_1}{T_1} = \frac{P_2}{T_2}
\]

\[
\frac{1 \text{ atm}}{298 \text{ K}} = \frac{x}{1474 \text{ K}}
\]

\[
(1474 \text{ K})(1 \text{ atm}) = (298 \text{ K})x
\]

\[
4.95 \text{ atm} = x
\]

- **The Combined Gas Law**

\[
\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}
\]

- No variable remains constant
- Temperature always in Kelvin Scale

**Example Problem:** The volume of a gas filled balloon is 30.0 L at 313 K and has a pressure of 153 kPa. What would the volume be at standard temperature and pressure (STP)?

\[
\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}
\]

\[
(153 \text{ kPa})(30.0 \text{ L}) = (101.325 \text{kPa})(x)
\]

\[
313 \text{ K} = 273 \text{ K}
\]

\[
39.5 \text{ L} = x
\]

Remember **Avogadro’s Law**
Gas Laws Practice:

1) A chemist collects 59.0 mL of sulfur dioxide gas on a day when the atmospheric pressure is 0.989 atm. On the next day, the pressure has changed to 0.967 atm. What will the volume of the SO$_2$ gas on the second day?

\[ P_1V_1 = P_2V_2 \]
\[ (0.989 \text{ atm})(0.0590 \text{ L}) = (0.967 \text{ atm}) x \]
\[ 0.0603422958 \text{ L} = x \]
\[ 0.0603 \text{ L} = x \]

2) A can contains a gas with a volume of 56 mL and 20.0 °C. What is the volume in the can if it is heated to 50.0 °C?

\[ \frac{V_1}{T_1} = \frac{V_2}{T_2} \]
\[ \frac{0.056 \text{ L}}{293 \text{ K}} = \frac{x}{323 \text{ K}} \]
\[ 0.0617337884 \text{ L} = x \]
\[ 0.0617 \text{ L} = x \]

3) A gas with a volume of 4.0L at a pressure of 90.0 kPa is allowed to expand until the pressure drops to 20.0 kPa. What is the new volume?

\[ P_1V_1 = P_2V_2 \]
\[ (90.0 \text{ kPa})(4.0L) = (20.0 \text{ kPa})x \]
\[ 18 \text{ L} = x \]

4) At a winter carnival, a balloon is filled with 5.00 L of helium at a temperature of 273 K. What will be the volume of the balloon when it is brought into a warm house at 295 K?

\[ \frac{V_1}{T_1} = \frac{V_2}{T_2} \]
\[ \frac{5.00 \text{ L}}{273 \text{ K}} = \frac{x}{295 \text{ K}} \]
\[ 5.40 \text{ L} = x \]
5) The initial temperature of a 1.00 liter sample of argon is 20.0° C. The pressure is decreased from 720 mm Hg to 360 mm Hg and the volume increases to 2.14 liters. What was the change in temperature of the argon?

\[
\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}
\]

\[
\frac{(720 \text{ mm Hg})(1.00 \text{ L})}{293 \text{ K}} = \frac{(360 \text{ mm Hg})(2.14 \text{ L})}{x}
\]

\[
x = 310 \text{ K}
\]

6) 2.2 L of hydrogen at 6.5 atm pressure is used to fill a balloon at a final pressure of 1.15 atm. What is its final volume?

\[
P_1V_1 = P_2V_2
\]

\[
(6.5 \text{ atm})(2.2 \text{ L}) = (1.15 \text{ atm})x
\]

\[
x = 12 \text{ L}
\]

7) The pressure in an automobile tire is 200. kPa at a temperature of 25°C. At the end of a journey on a hot sunny day the pressure has risen to 223 kPa. What is the temperature of the air in the tire?

\[
\frac{P_1}{T_1} = \frac{P_2}{T_2}
\]

\[
\frac{200. \text{ kPa}}{298 \text{ K}} = \frac{223 \text{ kPa}}{x}
\]

\[
x = 330 \text{ K}
\]

8) A sample of argon has a volume of 5.00 L and the pressure is 0.920 atm. If the final temperature is 30.0° C, the final volume is 5.7 L, and the final pressure is 800. mm Hg, what was the initial temperature of the argon?

\[
P_1V_1 = P_2V_2
\]

**must have the same units for pressure! Remember \(1 \text{ atm} = 760 \text{ mmHg}\)

\[
\frac{800 \text{ mmHg}}{760 \text{ mmHg}} = 1.05 \text{ atm}
\]

\[
\frac{(0.920 \text{ atm})(5.00 \text{ L})}{303 \text{ K}} = \frac{(1.05 \text{ atm})(5.7 \text{ L})}{x}
\]

\[
x = 230 \text{ K}
\]
• Atmospheric pressure is measured using a **Barometer**.

• The pressure of a gas is measured using a **Manometer**.

• As you increase in elevation the atmospheric pressure **decreases (Olympic Training in Denver)**.

  - A **Manometer** is a device to measure the pressure of an enclosed gas sample. A common simple manometer consists of a $U$ shaped tube of glass filled with some liquid. Typically the liquid is mercury because of its high density.
  - The height difference determines the pressure.

  $$\text{Gas pressure} = \text{atmospheric pressure} + h.$$  

Avogadro’s Law: under the same condition of **Temperature** and **Pressure**, equal **volumes** of all **gases** contain the same number of **particles**.
Calculate the pressure inside each flask, given an atmospheric pressure of 760 mmHg.

The gas in the flask has a higher pressure than 760 mmHg. The pressure is $760 \text{ mmHg} + 200 \text{ mmHg} = 960 \text{ mmHg}$.

The gas in the flask has a lower pressure than 760 mmHg. The pressure is $760 \text{ mmHg} - 350 \text{ mmHg} = 410 \text{ mmHg}$. 
1. The word *kinetic* comes from a Greek word that means “to move.” The *kinetic molecular theory* is based upon the assumption that particles of matter (atoms or molecules) are in constant *random motion*.

2. Of the three states of matter, which one has the most kinetic energy? **Gas**

3. Which state of matter has particles that are separated by the largest distance? **Gas**

4. A scientific theory is an explanation of some type of natural phenomena. Theories are normally developed from careful study of the way the world behaves. Let’s look at how gases behave and see if the kinetic molecular theory makes sense.

5. Compared to liquids and solids, gases tend to have **lower** densities. This can be explained because the particles of gas are in **constant motion**.

6. If you apply pressure to a sample of gas, it is fairly easy to **compress** its volume (think about what would happen to a balloon if you squeeze it gently). This can be explained because there is a lot of **empty space** between gas particles.

7. If someone sprays perfume in one corner of the room, eventually a person on the other side of the room can smell it. This can be explained because gas particles move **quick** and **random**. In general, we would expect lighter gas particles to travel **faster** than heavier gas particles.

8. When two gases mix together or move through each other, this process is known as **diffusion**. When gas particles escape out of a tiny hole in a container, this process is known as **effusion**. You should know the difference between these two words so you can avoid any confusion!

9. Kinetic molecular theory can be summarized as follows:
   
   a. Gas particles are in constant **motion**.
   b. Gas particles are separated by relatively **large** distances.
   c. When gas particles collide, they do not transfer kinetic energy.
   d. Gas particles have **no** attractive or repulsive forces between them.
   e. The kinetic energy of a gas is dependent on the **temperature** of the gas.
Ideal Gas Law and Dalton’s Law Notes

- **Ideal Gas Law**

\[ PV = nRT \]

R = 0.0821 \(\frac{atm \cdot L}{mol \cdot K}\) = 8.315 \(\frac{kPa \cdot L}{mol \cdot K}\) = 62.4 \(\frac{mmHg \cdot L}{mol \cdot K}\)

○ To determine which R value to use, look at your unit of____PRESSURE____

○ If gases follow the Kinetic Molecular Theory, they are said to be ideal gases.

○ Although there is no such thing as an ideal gas, the theory still provides a good model to explain gas properties (P,V,T)

- At 34°C, the pressure inside a nitrogen filled tennis ball with a volume of 0.148 L is 212 kPa. How many moles of nitrogen gas are in the tennis ball?

\[ PV = nRT \]

\((212 \text{ kPa})(0.148 \text{ L}) = n (8.315 \frac{kPa \cdot L}{mol \cdot K})(307 \text{ K})\)

0.0123 mol = n

- A deep underground cavern contains 2.24 x10⁶ L of methane gas (CH₄) at a pressure of 1.50x10³ kPa and a temperature of 315 K. How many kilograms of methane does the cavern contain?

First find moles then grams then kilograms

\[ PV = nRT \]

\((1.50x10³ \text{ kPa})(2.24 \times 10^6 \text{ L}) = n (8.315 \frac{kPa \cdot L}{mol \cdot K})(315 \text{ K})\)

n = 1.28 x 10⁶ mol

Now with moles convert to grams

\[
\frac{1.28 \times 10^6 \text{ mol of CH}_4}{1 \text{ mol of CH}_4} \times 16.05 \text{ g of CH}_4 = 2.06 \times 10^7 \text{ g of CH}_4
\]

Now with grams convert to kilograms

\[
\frac{2.06 \times 10^7 \text{ g of CH}_4}{1000 \text{ g of CH}_4} = 2.06 \times 10^4 \text{ kg of CH}_4
\]

**Ideal Gasses vs. Real Gasses**

**Ideal Gas assumptions:**
- No volume
- No attractive forces between particles of gas

There is no gas which exhibits these qualities, but there are many conditions which real gasses behave the same as an ideal gas.
Real gasses differ most from an ideal gas at low temperatures and high pressures.

- **Dalton’s Law of Partial Pressures**
  - The total pressure exerted by a mixture of gases is the sum of the individual of each gas.
  - Each individual gas behaves as if it were independent of the others.

  \[ P_{total} = P_1 + P_2 + P_3 + \ldots \]

  - For example, if two gases such as oxygen and nitrogen are present in a flask and \( P_{nitrogen} = 250 \text{ mmHg} \) and \( P_{oxygen} = 3.0 \times 10^2 \text{ mmHg} \), then the total pressure is 550 mmHg.

    \[
    P_{nitrogen} + P_{oxygen} = P_{total} \\
    250 \text{ mmHg} + 3.0 \times 10^2 \text{ mmHg} = 550 \text{ mmHg} \\
    250 \text{ mmHg} + 300 \text{ mmHg} = 550 \text{ mmHg}
    \]

- If CO\(_2\) gas is collected over water at 25 °C at atmospheric pressure (1.00 atm), and the partial pressure of water vapor at 25 °C is 0.54 atm. What is the partial pressure of the dry gas?

  \[
  P_{CO_2} + P_{H_2O} = P_{total} \\
  P_{CO_2} + 0.54 \text{ atm} = 1.00 \text{ atm} \\
  P_{CO_2} = 1.00 \text{ atm} - 0.54 \text{ atm} \\
  P_{CO_2} = 0.46 \text{ atm}
  \]

- Mole Fraction (X) – *the ratio of the number of moles of a given component in a mixture to the number of moles in the mixture.*
  - At constant volume and temperature, mole ratios and pressure ratios mean the same thing.

- The partial pressure of oxygen was observed to be 156 torr in the air with an atmospheric pressure of 743 torrs. Calculate the mole fraction of oxygen present.

  \[
  \text{number of moles} / \text{number of moles} = \text{Mole Fraction} \\
  156 \text{ torr} / 743 \text{ torrs} = 0.210
  \]
Ideal Gas Law and Dalton’s Law Practice:

1. **How many moles** of air molecules are contained in a 2.00 L flask at 98.9 atm and 245 K?

   \[
   \frac{PV}{RT} = \frac{(98.9 \text{ atm})(2.00 \text{ L})}{(0.0821 \text{ atm} \cdot \text{L}/\text{mol} \cdot \text{K})(245 \text{ K})}
   \]

   \[n = 9.83 \text{ mol}\]

2. How many moles of gas are contained in 51.41 liters at 101.325 kPa and 0°C?

   \[
   \frac{PV}{RT} = \frac{(101.325 \text{ kPa})(51.41 \text{ L})}{(8.315 \text{ kPa} \cdot \text{L}/\text{mol} \cdot \text{K})(273 \text{ K})}
   \]

   \[n = 2.295 \text{ mol}\]

3. How many moles of gases are contained in a can with a volume of 800.0 mL and a pressure of 600.0 atm at 20 °C?

   \[
   \frac{PV}{RT} = \frac{(600.0 \text{ atm})(0.8000 \text{ L})}{(0.0821 \text{ atm} \cdot \text{L}/\text{mol} \cdot \text{K})(293 \text{ K})}
   \]

   \[n = 19.95 \text{ mol}\]

4. Find the total pressure for a mixture that contains four gases with partial pressures of 0.39 kPa, 0.53 kPa, 0.73 kPa, 0.72 kPa.

   \[P_{\text{total}} = 0.39 \text{ kPa} + 0.53 \text{ kPa} + 0.73 \text{ kPa} + 0.72 \text{ kPa}\]

   \[P_{\text{total}} = 2.37 \text{ kPa}\]

5. Hydrogen gas is collected over water at 25°C; the atmospheric pressure and the partial pressure of water vapor at 25 °C is 3.17 kPa. What is the partial pressure of the hydrogen? Hint: if a reaction is conducted at atmospheric pressure, than that is the total pressure.

   (pressure of water vapor at 25 °C is 0.54 atm)

   \[P_{\text{total}} = P_{\text{H}_2} + P_{\text{H}_2\text{O}}\]

   \[3.17 \text{ kPa} = P_{\text{H}_2} + 0.54 \text{ kPa}\]

   \[P_{\text{H}_2} = 3.17 \text{ kPa} - 0.54 \text{ kPa}\]

   \[P_{\text{H}_2} = 2.63 \text{ kPa}\]
Unit 9 Review: Gas Laws

See Separate Key