Chapter 11: Gases:

**Homework:** Read Chapter 11. Work out sample/practice exercises

**Bonus**  Ch 11: 29, 33, 39, 45, 53, 61, 69, 73, 77, 81, 85, 87, 91, 93, 97, 105
Check MasteringChemistry deadlines

**Gas Properties:**

Gases have …

- high kinetic energy
- low density (measured in g/L)
- homogeneous mixing
- gases take on the volume and shape of its container, filling the container completely and evenly
- gases exert pressure
- gases are compressible and expandable.

**Kinetic Molecular Theory (KMT) of Gases:**

*KMT explains behavior of gases and predicts the correct behavior for ideal gases.*

Assumptions for the KMT…

1) An ideal gas consists of tiny particles in constant, random, straight-line motion.
2) Particles exhibit no attraction or repulsion between each other. Collisions of gas particles are elastic. *Assumes gas particles will not stick together.*
3) There is a lot of space between particles. The volume of the particles themselves is negligible compared to the total volume of the gas. *Assumes an ideal gas takes up no volume - point mass.*
4) The speed of particles increases with increasing temperature. The average kinetic energy of gas particles is proportional to the Temperature in Kelvin.

*This model is not perfect, it begins to break down when temperatures are lower (gas particles begin to stick and become liquid) and pressures are higher (with many particles that cause the higher pressures, the gas particles start occupying some of the volume of the container).*
Ideal Gas:

An ideal gas consists of tiny particles in constant, random, straight-line motion. Particles exhibit no attraction or repulsion between each other. Collisions of gas particles are elastic (no energy, friction or momentum lost, just transferred). The volume of the particles themselves is negligible compared to the total volume of the gas. The total kinetic energy is constant at constant T. The average kinetic energy of gas particles is proportional to the Temperature in Kelvin.

When gas is made up of a mixture of species, like air, each component of gas acts independently of the others.

Gases are not always acting ideally:

At lowered temperatures the gas particles begin to stick together and start to change phase to a liquid. At higher pressures the many gas particles start occupying some of the volume of the container.
Gas Pressure:

- Pressure is the force exerted per unit area by gas molecules as they collide with the surfaces around them.
- Because of pressure, we can drink from straws, inflate basketballs, and move air into and out of our lungs.
- Variation in pressure in Earth’s atmosphere creates wind, and changes in pressure help predict weather. Pressure is all around us and even inside us.
- A molecule exerts a force when it collides with a surface. The result of many of these collisions is pressure.
- On Earth at sea level, the gas molecules in our atmosphere exert an average pressure of 101,325 N/m$^2$ or Pascals, in English units, 14.7 lb/in$^2$ (psi), common units, 1 atmosphere and 760 torr or mm Hg

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

Pressure is Force per unit Area…

\[
P = \frac{\text{mass} \times \text{acceleration}}{\text{Area}}
\]

The SI unit for pressure is Pascal is very small.

\[
1 \text{ Pa} = \frac{\text{Pressure in kg} \times \text{acceleration in m/s}^2}{\text{area in m}^2} = \text{N/m}^2
\]
Pressure:

- Pressure on a mountaintop is less than sea level and sea level is less than low places like Death Valley.

- As we climb a mountain or ascend in an airplane, there are fewer molecules per unit volume in air and the pressure drops. Most commercial airplanes fly at elevations between 25,000 and 40,000 ft, where atmospheric pressure is below 0.50 atm. The physiological effects of these lowered pressures—and the correspondingly lowered oxygen levels—include dizziness, headache, shortness of breath, and even unconsciousness. Federal regulations require cabin pressure in commercial airliners be pressurized greater than the equivalent of outside air pressure at 8000 ft. which is 0.72 atm.

- You may feel the effect of a drop in pressure as a pain in your ears. The external pressure drops while the pressure of the air within your ear cavities remains the same. This creates an imbalance—the lower external pressure causes your eardrum to bulge outward, causing pain.

- With time, the excess air within your ears’ cavities escapes, equalizing the internal and external pressure and relieving the pain.

Pressure Units:

We commonly use units of **atmospheres** or **torr**. Evangelista Torricelli (1608-1647) based pressure measurements on a barometer filled with liquid mercury.

\[
1 \text{ atm} = 760 \text{ torr or 760 mm Hg}
\]

(remember this conversion)

\[
1 \text{ atm} = 101,325 \text{ Pa or 101.325 kPa or 1.01325 bar}
\]

\[
1 \text{ atm} = 14.7 \text{ psi}
\]

\[
1 \text{ atm} = 29.92 \text{ in. Hg}
\]

\[
1 \text{ bar} = 0.986923 \text{ atm}
\]
**The Gas Laws:** All the gas laws apply to all pure gases, regardless of chemical identity or mixtures.

**Boyle’s Law (1662):** \( PV = k \) or \( P_1V_1 = P_2V_2 \)

**Charle’s Law (1787):** \( \frac{V}{T} = \text{constant} \) or \( \frac{V_1}{T_1} = \frac{V_2}{T_2} \) (must use Kelvin!)
- Heating the air in a balloon makes it expand (Charles’s law).

- As the volume occupied by the hot air increases, its density decreases, allowing the balloon to float in the cooler, denser air that surrounds it.

- If temperature is changed when a balloon is moved from an ice-water bath into a boiling-water bath, the gas molecules inside it move faster due to the increased temperature. If the external pressure remains constant, the volume will change when the molecules expand the balloon and collectively occupy a larger volume.

**Gay-Lussac’s Law of Combining Volumes (1808):** At constant temperature and pressure, the volumes of gases involved in chemical reactions are in ratios of small whole numbers.

\[ \text{N}_2 \,(g) + 3 \, \text{H}_2 \,(g) \rightarrow 2 \, \text{NH}_3 \,(g) \]

Each gas species must be at the same temperature and partial pressures.

1 volume of \( \text{N}_2 \) + 3 volume of \( \text{H}_2 \) = 2 volumes of \( \text{NH}_3 \) (all gases)

5.0 L of \( \text{N}_2 \) react with _____L of \( \text{H}_2 \) to produce _____L of \( \text{NH}_3 \)
Avogadro’s Law (1810):
\[ \frac{V}{n} = \text{constant} \]
or
\[ \frac{V_1}{n_1} = \frac{V_2}{n_2} \]

Ideal Gas Law (first stated by Clapeyron in 1834):
\[ PV = nRT, \quad \text{where } R = 0.08206 \text{ L·atm/mol·K} \]

Combined Gas Law: \[ P_1 \frac{V_1}{n_1} T_1 = P_2 \frac{V_2}{n_2} T_2 \] or \[ P_1 \frac{V_1}{T_1} = P_2 \frac{V_2}{T_2} \]

Variations of Ideal Gas Law:
\[ MW = \frac{\text{mass}RT}{PV} \]
\[ \text{Density} = \frac{\text{(MW)P}}{RT} \]

### TABLE 11.2 Relationships between Simple Gas Laws and Ideal Gas Law

<table>
<thead>
<tr>
<th>Variable Quantities</th>
<th>Constant Quantities</th>
<th>Ideal Gas Law in Form of Variables-constant</th>
<th>Simple Gas Law</th>
<th>Name of Simple Law</th>
</tr>
</thead>
<tbody>
<tr>
<td>(V) and (P)</td>
<td>(n) and (T)</td>
<td>(PV = nRT)</td>
<td>(P_1 V_1 = P_2 V_2)</td>
<td>Boyle’s law</td>
</tr>
<tr>
<td>(V) and (T)</td>
<td>(n) and (P)</td>
<td>(\frac{V}{T} = \frac{nR}{P})</td>
<td>(\frac{V_1}{T_1} = \frac{V_2}{T_2})</td>
<td>Charles’s law</td>
</tr>
<tr>
<td>(P) and (T)</td>
<td>(n) and (V)</td>
<td>(\frac{P}{T} = \frac{nR}{V})</td>
<td>(\frac{P_1}{T_1} = \frac{P_2}{T_2})</td>
<td>Gay-Lussac’s law</td>
</tr>
<tr>
<td>(P) and (n)</td>
<td>(V) and (T)</td>
<td>(\frac{P}{n} = \frac{RT}{V})</td>
<td>(\frac{P_1}{n_1} = \frac{P_2}{n_2})</td>
<td></td>
</tr>
<tr>
<td>(V) and (n)</td>
<td>(T) and (P)</td>
<td>(\frac{V}{n} = \frac{RT}{P})</td>
<td>(\frac{V_1}{n_1} = \frac{V_2}{n_2})</td>
<td>Avogadro’s law</td>
</tr>
</tbody>
</table>
STP:
Standard Temperature and Pressure is at 0°C or 273.15K and 1 atm or 760 torr
Molar volume under STP conditions for an ideal gas = 22.4 L/mol
Note: recently scientists have redefined STP as a pressure of 1 bar in place of 1 atm. By this new definition the standard molar volume is 22.7 L/mol. We will continue to use the old standard for this class.
Personally, I wish they would redefine the STP temperature to 25°C since that is a more practical room temperature for experiments.

Partial Pressure:
- Each of the components in a gas mixture acts independently of the others.
- The pressure due to any individual component in a gas mixture is called the partial pressure of that component.
- The air in our atmosphere is a mixture containing 78% nitrogen, 21% oxygen, 0.9% argon, 0.04% carbon dioxide, and a few other gases in smaller amounts.
- The oxygen molecules in air exert a certain pressure—21% of the total pressure—that is also independent of the presence of the other gases in the mixture.

Dalton’s Law of Partial Pressure (1801):
\[ P_{\text{total}} = P_1 + P_2 + P_3 \ldots \text{ at constant } V \text{ and } T \text{ (same container)} \]
This law of partial pressures is especially important when collecting gas by displacement of water.
Collecting Gases over Water:

To calculate the partial pressure of an experimentally collected gas by displacement of water one must subtract the partial pressure due to the vapor pressure of water in the closed container. Vapor pressure is dependent on temperature and can be found tabulated in tables and reference books.

Vapor Pressure:

All closed containers that are partially filled with liquid will have some particles of that liquid in the gas state above the liquid level. Any gas collected by water displacement will be contaminated with water vapor. This amount depends on the temperature.

<table>
<thead>
<tr>
<th>Temperature °C</th>
<th>Pressure (mmHg)</th>
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<th>Pressure (mmHg)</th>
</tr>
</thead>
<tbody>
<tr>
<td>0</td>
<td>4.58</td>
<td>30</td>
<td>31.9</td>
</tr>
<tr>
<td>5</td>
<td>6.54</td>
<td>40</td>
<td>55.4</td>
</tr>
<tr>
<td>10</td>
<td>9.21</td>
<td>50</td>
<td>92.6</td>
</tr>
<tr>
<td>15</td>
<td>12.79</td>
<td>75</td>
<td>289.1</td>
</tr>
<tr>
<td>20</td>
<td>17.55</td>
<td>100</td>
<td>760.0</td>
</tr>
<tr>
<td>25</td>
<td>23.78</td>
<td>110</td>
<td>1075</td>
</tr>
</tbody>
</table>

*Note: The vapor pressure of water at its normal boiling point is equivalent to 1 atm

Gas Examples:

1. Convert 748 torr to the following units…
   a) atm, b) Pa, c) bar, d) psi
2. Every 10 m depth under water adds 1 atm of pressure. At the surface of the water the pressure is 1.0 atm. A scuba diver descends to 20 m and the pressure is now 3.0 atm. The regulator must deliver the air at a pressure of 3.0 atm for the diver to breathe. What would happen to the lungs of a diver if he holds his breath while rising from a depth of 20 m to the surface?
3. What is the number of moles of gas in a 550-ml container at 30.0°C and 751 torr pressure?

4. An expandable container under constant pressure may expand to a maximum of 10.0L. If the volume of the container is initially 250 ml at 18°C, what is the maximum temperature in Celsius the container may be heated at constant pressure?

5. Helium in a tank sold for filling party balloons is at a pressure of 18.0 atm at 22°C. The tank has a volume of 100 L. How many balloons can be filled from a tank if the balloons have a volume of 2.3 L at 22°C and the pressure inside the balloons is 1.00 atm? (Note: 1 atm of pressure must remain in the tank, it will not be emptied to a vacuum)

6. Calculate the moles, molecular weight & the density (g/L) of an unknown gas, if the gas sample weighs 0.222 g in a 348 mL container at 26.0 °C & 742-torr? Is this gas CH₄ or O₂?

7. Hydrogen gas is produced by the reaction of sodium metal with an excess of hydrochloric acid solution. The hydrogen gas was collected by water displacement at 22°C in a 127 ml container with a total pressure of 748 torr. The vapor pressure of water at 22°C is 19.8 torr.
   a) What mass of sodium metal was consumed in the reaction?
   b) What is the volume of dry H₂ gas at STP?

8. A 4.50 liter helium balloon is given to a small child by the sea. It is a nice warm day of 30°C and 1 atm. pressure. Unhappily, the child lets go of the balloon and it rises up into the air. At a new higher altitude the pressure is 703 torr and Temperature is 10°C. What is the new volume of the balloon?

9. The density of a gas at STP is 0.7143 g/L. Is this gas CH₄, Ar, CO₂, or Cl₂?

10. What is the volume of 32.7 grams of F₂ gas at STP?

11. 534-mL of CH₄ gas was collected at 23°C and 754 torr. What new volume would the CH₄ gas occupy at 732 torr and 46°C?

12. 127 ml of N₂ gas was collected by displacement of water at 25°C and 742 torr. What is the number of moles of N₂ gas? (The vapor pressure of water at 25°C is 24 torr.)

13. What pressure will 4.15 g of CO₂ gas have when placed in a 320-mL container at 45.0°C?
14. Nitrogen gas is placed into a steel tank until its pressure is 21.4 atm at 20°C. The maximum pressure the tank can safely contain is 50.0 atm. What is the highest temperature in °C that this tank be safely heated.

15. a) Balance the following unbalanced equation:
   \[ \text{C}_5\text{H}_12 (g) \quad + \quad \text{O}_2 (g) \quad \rightarrow \quad \text{CO}_2 (g) \quad + \quad \text{H}_2\text{O} (g) \]

   b) What volume of CO\(_2\) gas at STP will be produced from 100 g of C\(_5\)H\(_12\)?

16. Balance \[ \text{H}_2 (g) \quad + \quad \text{N}_2 (g) \quad \rightarrow \quad \text{NH}_3 (g) \] all gases at same temperature and pressure. How many liters of H\(_2\) gas are required to produce 200 liters of ammonia gas?

17. A 2.135-g gas sample occupies 315.5 ml at 739.2 mm Hg and 26.1°C. Analysis of the compound shows it to be 14.05% C, 41.48% Cl, and 44.46% F by mass. What is the molecular formula?

18. Typically, when a person coughs, he or she first inhales about 2.0-L of air at 1.0-atm pressure and 25°C. The epiglottis and the vocal cords then shut, trapping the air in the lungs, where it is warmed to 37°C and compressed to 1.7-L by action of the diaphragm and chest muscles. The sudden opening of the epiglottis and vocal cords releases this air explosively. What is the approximate pressure of the gas inside the lungs just before the cough?

19. Which of the following occupies the greatest volume when measure at STP?
   a) 30.0 g O\(_2\) b) 1.10 mol SO\(_2\) c) 20.0 L CO d) 7.2 x 10\(^{23}\) molecules Cl\(_2\)

20. Carbon dioxide gas (CO\(_2\)) is collected over water in a 544 ml container at 23.0°C and equalized with the room pressure of 750 torr. The vapor pressure of water at 23.0°C is 21.1 torr. What volume will this same sample of dry CO\(_2\) gas occupy under STP conditions?

**Gas Answers:**

1. a) 0.984 atm, b) 9.97 x 10\(^4\) Pa, c) 0.997 bar, d) 14.5 psi

2. The lungs would expand by a factor of 3 times and burst, Yikes! for the unfortunate diver. \[ P_1V_1 = P_2V_2 \]

3. 0.0218 mol

4. \(11640 \text{ K or 11370 °C} \)

5. Around 739 balloons
6. 0.0138 moles, molecular weight 16.0 g/mol & the density 0.638 g/L; unknown gas is CH₄

7. a) 2 Na + 2 HCl → 2 NaCl + H₂

   728.2 torr = 0.958 atm H₂, Moles of H₂ = 5.03 x 10⁻³ moles, **0.231 g Na**

   b) **0.113 L dry H₂**

8. 4.54 L

9. CH₄

10. 19.3 L

11. 593 ml

12. 4.90 x 10⁻³ mol

13. 7.69 atm

14. 411°C

15. a) C₅H₁₂ (g) + 8 O₂ (g) → 5 CO₂ (g) + 6 H₂O (g)

   b) 156 L

16. 3 H₂ (g) + N₂ (g) → 2 NH₃ (g) 300 liters of H₂

17. 170.9 g/mol, C₂Cl₂F₄

18. 1.22 atm

19. d) 7.2 x 10⁻³ molecules Cl₂

20. 0.481 L