The Ideal Gas Law is very simply expressed:

\[ PV = nRT \]

from which simpler gas laws such as Boyle's, Charles's, Avogadro's and Amonton's law be derived.

**Introduction**

Many chemists had dreamed of having an equation that describes relation of a gas molecule to its environment such as pressure or temperature. However, they had encountered many difficulties because of the fact that there always are other affecting factors such as intermolecular forces. Despite this fact, chemists came up with a simple gas equation to study gas behavior while putting a blind eye to minor factors.

When dealing with gas, a famous equation was used to relate all of the factors needed in order to solve a gas problem. This equation is known as the Ideal Gas Equation. As we have always known, anything ideal does not exist. In this issue, two well-known assumptions should have been made beforehand:

1. the particles have no forces acting among them, and
2. these particles do not take up any space, meaning their atomic volume is completely ignored.

An **ideal gas** is a hypothetical gas dreamed by chemists and students because it would be much easier if things like intermolecular forces do not exist to complicate the simple *Ideal Gas Law*. Ideal gases are essentially point masses moving in constant, random, straight-line motion. Its behavior is described by the assumptions listed in the [Kinetic-Molecular Theory of Gases](#). This definition of an ideal gas contrasts with the Non-Ideal Gas definition, because this equation represents how gas actually behaves in reality. For now, let us focus on the Ideal Gas.

We must emphasize that this gas law is *ideal*. As students, professors, and chemists, we sometimes need to understand the concepts before we can apply it, and assuming the gases are in an ideal state where it is unaffected by real world conditions will help us better understand the behavior the gases. In order for a gas to be *ideal*, its behavior must follow the [Kinetic-Molecular Theory](#) whereas the Non-Ideal Gases will deviate from this theory due to real world conditions.

**The Ideal Gas Equation**

Before we look at the **Ideal Gas Equation**, let us state the four gas variables and one constant for a better understanding. The four gas variables are: pressure (P), volume (V), number of mole of gas (n), and temperature (T). Lastly, the constant in the equation shown below is R, known as the the *gas constant*, which will be discussed in depth further later:

\[ PV = nRT \]

Another way to describe an ideal gas is to describe it in mathematically. Consider the following equation:

\[ \frac{PV}{nRT} = 1 \]

The term \( \frac{PV}{nRT} \) is also called the *compression factor* and is a measure of the ideality of the gas. An ideal gas
will always equal 1 when plugged into this equation. The greater it deviates from the number 1, the more it will behave like a real gas rather than an ideal. A few things should always be kept in mind when working with this equation, as you may find it extremely helpful when checking your answer after working out a gas problem.

- Pressure is directly proportional to number of molecule and temperature. (Since P is on the opposite side of the equation to $n$ and $T$)
- Pressure, however, is indirectly proportional to volume. (Since P is on the same side of the equation with V)

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**Simple Gas Laws**

The Ideal Gas Law is simply the combination of all Simple Gas Laws (Boyle's Law, Charles' Law, and Avogadro's Law), and so learning this one means that you have learned them all. The Simple Gas Laws can always be derived from the Ideal Gas equation.

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**Boyle's Law**

Boyle's Law describes the inverse proportional relationship between pressure and volume at a constant temperature and a fixed amount of gas. This law came from a manipulation of the Ideal Gas Law.

$$ P \propto \dfrac{1}{V} $$

or expressed from two pressure/volume points:

$$ P_1V_1 = P_2V_2 $$

This equation would be ideal when working with problem asking for the initial or final value of pressure or volume of a certain gas when one of the two factor is missing.
Charles's Law

Charles's Law describes the directly proportional relationship between the volume and temperature (in Kelvin) of a fixed amount of gas, when the pressure is held constant.

\[ V \propto \; T \]

or express from two volume/temperature points:

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

This equation can be used to solve for initial or final value of volume or temperature under the given condition that pressure and the number of mole of the gas stay the same.

Avogadro's Law

Volume of a gas is directly proportional to the amount of gas at a constant temperature and pressure.

\[ V \propto \; n \]

or expressed as a two volume/number points:

\[ \frac{V_1}{n_1} = \frac{V_2}{n_2} \]

Avogadro's Law can apply well to problems using Standard Temperature and Pressure (see below), because of a set amount of pressure and temperature.

Amontons's Law

Given a constant number of mole of a gas and an unchanged volume, pressure is directly proportional to temperature.

\[ P \propto \; T \]

or expressed as two pressure/temperature points:

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

Boyle's Law, Charles' Law, and Avogadro's Law and Amontons's Law are given under certain conditions so directly combining them will not work. Through advanced mathematics (provided in outside link if you are interested), the properties of the three simple gas laws will give you the Ideal Gas Equation.

Standard Temperature and Pressure (STP)

Standard condition of temperature and pressure is known as STP. Two things you should know about this is listed below.

- The universal value of STP is 1 atm (pressure) and 0°C. Note that this form specifically stated 0°C degree, not 273...
Kelvin, even though you will have to convert into Kelvin when plugging this value into the Ideal Gas equation or any of the simple gas equations.

- In STP, 1 mole of gas will take up 22.4 L of the volume of the container.

Units of P, V and T

The table below lists the different units for each property.

<table>
<thead>
<tr>
<th>Factor</th>
<th>Variable</th>
<th>Units</th>
</tr>
</thead>
<tbody>
<tr>
<td>Pressure</td>
<td>P</td>
<td>atm, Torr, Pa, mmHg</td>
</tr>
<tr>
<td>Volume</td>
<td>V</td>
<td>L, m³</td>
</tr>
<tr>
<td>Moles</td>
<td>n</td>
<td>mol</td>
</tr>
<tr>
<td>Temperature</td>
<td>T</td>
<td>K</td>
</tr>
<tr>
<td>Gas Constant</td>
<td>R*</td>
<td>see Values of R table below</td>
</tr>
</tbody>
</table>

Take note of certain things such as temperature is always in its SI units of Kelvin (K) rather than Celsius (C), and the amount of gas is always measured in moles. Gas pressure and volume, on the other hand, may have various different units, so be sure to know how to convert to the appropriate units if necessary.

Pressure Units

Use the following table as a reference for pressure.

*Common Units of Pressure*

<table>
<thead>
<tr>
<th>Unit</th>
<th>Symbol</th>
<th>Equivalent to 1 atm</th>
</tr>
</thead>
<tbody>
<tr>
<td>Atmosphere</td>
<td>atm</td>
<td>1 atm</td>
</tr>
<tr>
<td>Millimeter of Mercury</td>
<td>mmHg</td>
<td>760 mmHg</td>
</tr>
<tr>
<td>Torr</td>
<td>Torr</td>
<td>760 Torr</td>
</tr>
<tr>
<td>Pascal</td>
<td>Pa</td>
<td>101326 Pa</td>
</tr>
<tr>
<td>Kilopascal</td>
<td>kPa*</td>
<td>101.326 kPa</td>
</tr>
<tr>
<td>Bar</td>
<td>bar</td>
<td>1.01325 bar</td>
</tr>
<tr>
<td>Millibar</td>
<td>mb</td>
<td>1013.25 mb</td>
</tr>
</tbody>
</table>
The Gas Constant (R)

Here comes the tricky part when it comes to the gas constant, R. Value of R WILL change when dealing with different unit of pressure and volume (Temperature factor is overlooked because temperature will always be in Kelvin instead of Celsius when using the Ideal Gas equation). Only through appropriate value of R will you get the correct answer of the problem. It is simply a constant, and the different values of R correlates accordingly with the units given. When choosing a value of R, choose the one with the appropriate units of the given information (sometimes given units must be converted accordingly). Here are some commonly used values of R:

<table>
<thead>
<tr>
<th>Values of R</th>
</tr>
</thead>
<tbody>
<tr>
<td>0.082057 L atm mol⁻¹ K⁻¹</td>
</tr>
<tr>
<td>62.364 L Torr mol⁻¹ K⁻¹</td>
</tr>
<tr>
<td>8.3145 m³ Pa mol⁻¹ K⁻¹</td>
</tr>
<tr>
<td>8.3145 J mol⁻¹ K⁻¹</td>
</tr>
</tbody>
</table>

*note: This is the SI unit for the gas constant

Example 1

So, which value of R should I use?

Solution

Because of the various value of R you can use to solve a problem. It is crucial to match your units of Pressure, Volume, number of mole, and Temperature with the units of R.

- If you use the first value of R, which is 0.082057 L atm mol⁻¹ K⁻¹, your unit for pressure must be atm, for volume must be liter, for temperature must be Kelvin.
- If you use the second value of R, which is 62.364 L Torr mol⁻¹ K⁻¹, your unit for pressure must be Torr, for volume must be liter, and for temperature must be Kelvin.

Ideal Gas Law Applications

How do you know the Ideal Gas Equation is the correct equation to use? Use the Ideal Gas Equation to solve a problem when the amount of gas is given and the mass of the gas is constant. There are various type of problems that will require the use of the Ideal Gas Equation.

- Solving for the unknown variable
• Initial and Final
• Partial Pressure

Other things to keep in mind: Know what Standard Temperature and Pressure (STP) values are. Know how to do Stoichiometry. Know your basic equations. Take a look at the problems below for examples of each different type of problem. Attempt them initially, and if help is needed, the solutions are right below them. Remark: The units must cancel out to get the appropriate unit; knowing this will help you double check your answer.

Example 2

5.0 g of neon is at 256 mm Hg and at a temperature of 35º C. What is the volume?

**SOLUTION**

**Step 1: Write down your given information:**

- \( P = 256 \text{ mmHg} \)
- \( V = ? \)
- \( m = 5.0 \text{ g} \)
- \( R = 0.0820574 \text{ L·atm·mol}^{-1}·\text{K}^{-1} \)
- \( T = 35º \text{ C} \)

**Step 2: Convert as necessary:**

Pressure: \( 256 \text{ mmHg} \times \left( \frac{1 \text{ atm}}{760 \text{ mmHg}} \right) = 0.3368 \text{ atm} \)

Moles: \( 5.0 \text{ g} \text{ Ne} \times \left( \frac{1 \text{ mol}}{20.1797 \text{ g}} \right) = 0.25 \text{ mol} \text{ Ne} \)

Temperature: \( 35º \text{ C} + 273 = 308 \text{ K} \)

**Step 3: Plug in the variables into the appropriate equation.**

\[
V = \left( \frac{nRT}{P} \right)
\]

\[
V = \left( \frac{0.25 \text{ mol} \times 0.0820574 \text{ L·atm·mol}^{-1}·\text{K}^{-1} \times 308 \text{ K}}{0.3368 \text{ atm}} \right)
\]

\[
V = 19 \text{ L}
\]

Example 3

What is a gas’s temperature in Celsius when it has a volume of 25 L, 203 mol, 143.5 atm?

**SOLUTION**

**Step 1: Write down your given information:**

- \( P = 143.5 \text{ atm} \)
Step 2: *Skip because all units are the appropriate units.*

Step 3: *Plug in the variables into the appropriate equation.*

\[
T = \frac{PV}{nR}
\]

\[
T = \frac{(143.5 \text{ atm})(25 \text{ L})}{(203 \text{ mol})(0.08206 \text{ L\cdot atm/K mol})}
\]

\[
T = 215.4 \text{ K}
\]

Step 4: *You are not done. Be sure to read the problem carefully, and answer what they are asking for. In this case, they are asking for temperature in Celsius, so you will need to convert it from K, the units you have.*

\[
215.4 \text{ K} - 273 = -57.4 ^\circ \text{ C}
\]

Example 4

What is the density of nitrogen gas \((N_2)\) at 248.0 Torr and 18\(^\circ\) C?

**SOLUTION**

**Step 1: Write down your given information**

- \(P = 248.0\) Torr
- \(V = ?\)
- \(n = ?\)
- \(R = 0.0820574 \text{ L\cdot atm\cdot mol}^{-1} \text{ K}^{-1}\)
- \(T = 18 ^\circ \text{ C}\)

**Step 2: Convert as necessary.**

\[
(248 \text{ Torr}) \times \frac{1 \text{ atm}}{760 \text{ Torr}} = 0.3263 \text{ atm}
\]

\[
18 ^\circ \text{ C} + 273 = 291 \text{ K}
\]

**Step 3: This one is tricky. We need to manipulate the Ideal Gas Equation to incorporate density into the equation.** *Write down all known equations:

\[PV = nRT\]

\[
\rho = \frac{m}{V}
\]
where \(\rho=\text{density}, \ m=\text{mass}, \ V=\text{Volume}\)

\[m=M \times n\]

where \(m=\text{mass}, \ M=\text{molar mass}, \ n=\text{moles}\)

*Now take the density equation.

\[\rho=\dfrac{m}{V}\]

*Keeping in mind \((m=M \times n)\)...replace \((m=M \times n)\) for \(m\) within the density formula.

\[\rho=\dfrac{M \times n}{V}\]

\[\dfrac{\rho}{M} = \dfrac{n}{V}\]

*Now manipulate the Ideal Gas Equation

\((PV = nRT)\)

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

\[\dfrac{n}{V}\] is in both equations.

\[\dfrac{n}{V} = \dfrac{\rho}{M}\]

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

*Now combine them please.

\[\dfrac{\rho}{M} = \dfrac{P}{RT}\]

*Isolate density.

\[\rho = \dfrac{PM}{RT}\]

**Step 4: Now plug in the information you have.**

\[\rho = \dfrac{PM}{RT}\]

\[\rho = \dfrac{(0.3263 \ \text{atm})(2\times14.01 \ \text{g/mol})}{(0.08206 \ L \text{ atm/K mol})(291 \ \text{K})}\]

\[\rho = 0.3828 \ \text{g/L}\]

Example 5

Find the volume, in mL, when 7.00 g of \(\text{O}_2\) and 1.50 g of \(\text{Cl}_2\) are mixed in a container with a pressure of 482 atm and at a temperature of 22º C.
**SOLUTION**

**Step 1:** Write down your given information

- \( P = 482 \text{ atm} \)
- \( V = ? \)
- \( n = ? \)
- \( R = 0.0820574 \text{ L} \cdot \text{atm} \cdot \text{mol}^{-1} \cdot \text{K}^{-1} \)
- \( T = 22^\circ \text{ C} + 273 = 295 \text{ K} \)
- 1.50g \( \text{Cl}_2 \)
- 7.00g \( \text{O}_2 \)

**Step 2:** Find the total moles of the mixed gases in order to use the Ideal Gas Equation.

\[
[n_{\text{total}} = n_{\text{O}_2} + n_{\text{Cl}_2}]
\]

\[
= \left[7.0 \text{ g} \text{ O}_2 \times \frac{1 \text{ mol} \text{ O}_2}{32.00 \text{ g} \text{ O}_2}\right] + \left[1.5 \text{ g} \text{ Cl}_2 \times \frac{1 \text{ mol} \text{ Cl}_2}{70.905 \text{ g} \text{ Cl}_2}\right]
\]

\[
= 0.2188 \text{ mol} \text{ O}_2 + 0.0212 \text{ mol} \text{ Cl}_2
\]

\[
= 0.24 \text{ mol}
\]

**Step 3:** Now that you have moles, plug in your information in the Ideal Gas Equation.

\[
V = \frac{nRT}{P}
\]

\[
V = \frac{(0.24 \text{ mol})(0.08206 \text{ L atm/K mol})(295 \text{ K})}{(482 \text{ atm})}
\]

\[
V = 0.0121 \text{ L}
\]

**Step 4:** Almost done! Now just convert the liters to milliliters.

\[
0.0121 \text{ L} \times \frac{1000 \text{ mL}}{1 \text{ L}} = 12.1 \text{ mL}
\]

**Example 6**

A 3.00 L container is filled with \( \text{Ne}_g \) at 770 mmHg at 27^\circ \text{ C}. A \( 0.633 \text{ g} \text{ CO}_2 \) vapor is then added. What is the partial pressure of \( \text{CO}_2 \) and \( \text{Ne} \) in atm? What is the total pressure in the container in atm?

**SOLUTION**

**Step 1:** Write down all given information, and convert as necessary.

Before:

- \( P = 770 \text{ mmHg} \rightarrow 1.01 \text{ atm} \)
• $V = 3.00\text{L}$
• $n_{\text{Ne}} =$
• $T = 27^\circ\text{C} \rightarrow (300\text{K})$

Other Unknowns: $(n_{\text{(CO}_2\text{)}})$=

\[n_{\text{(CO}_2\text{)}}} = 0.633\text{g} \cdot \dfrac{1\text{mol}}{44\text{g}} = 0.0144\text{mol} \cdot \text{CO}_2\]

\[n_{\text{Ne}} = \dfrac{PV}{RT}\]
\[n_{\text{Ne}} = \dfrac{(1.01\text{atm})(3.00\text{L})}{(0.08206\text{atm}\cdot\text{L/mol}\cdot\text{K})(300\text{K})}\]
\[n_{\text{Ne}} = 0.123\text{mol}\]

Because the pressure of the container before the $(\text{CO}_2)$ was added contained only $(\text{Ne})$, that is your partial pressure of $(\text{Ne})$. After converting it to atm, you have already answered part of the question!

\[P_{\text{Ne}} = 1.01\text{atm}\]

Step 3: Now that have pressure for $\text{Ne}$, you must find the partial pressure for $(\text{CO}_2)$. Use the ideal gas equation.

\[\dfrac{P_{\text{Ne}}V}{n_{\text{Ne}RT}} = \dfrac{P_{\text{CO}_2}V}{n_{\text{CO}_2}RT}\]

but because both gases share the same Volume $(V)$ and Temperature $(T)$ and since the Gas Constant $(R)$ is constants, all three terms cancel and can be removed them from the equation.

\[\dfrac{P_{\text{Ne}}}{n_{\text{Ne}}} = \dfrac{P_{\text{CO}_2}}{n_{\text{CO}_2}}\]
\[\dfrac{1.01\text{atm}}{0.123\text{mol}} = \dfrac{P_{\text{CO}_2}}{0.0144\text{mol}}\]
\[P_{\text{CO}_2} = 0.118\text{atm}\]

This is the partial pressure $(\text{CO}_2)$.

\[P_{\text{total}} = P_{\text{Ne}} + P_{\text{CO}_2}\]
\[P_{\text{total}} = 1.01\text{atm} + 0.118\text{atm}\]
\[P_{\text{total}} = 1.128\text{atm} \approx 1.13\text{atm} \text{ (with appropriate significant figures)}\]

References


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**Contributors**

- Duke LeTran (UCD)