Mathematics-based Chemistry for Science, engineering & preprofessional majors
Knowledge is Understanding

People “learn” in different ways BUT

1. They must be **self motivated**
2. They must be actively involved
3. Knowledge results ONLY through Active Participation
WHAT IS CHEMISTRY?

Chemistry is the Study Of _________ & the ______________________

- **Matter** is defined as anything that has both ______________________
Gas
Total disorder; much empty space; particles have complete freedom of motion; particles far apart

Cool or compress
Heat or reduce pressure

Liquid
Disorder; particles or clusters of particles are free to move relative to each other; particles close together

Cool
Heat

Crystalline solid
Ordered arrangement; particles are essentially in fixed positions; particles close together
Chapter 10

THE STUDY OF GASES

1. What are the fundamental properties of the gases that compose the air around us?

2. What LAWS describe their behavior?

3. What kind of THEORY can explain these properties and laws?
<table>
<thead>
<tr>
<th>Formula</th>
<th>Name</th>
<th>Characteristics</th>
</tr>
</thead>
<tbody>
<tr>
<td>HCN</td>
<td>Hydrogen cyanide</td>
<td>Very toxic, slight odor of bitter almonds</td>
</tr>
<tr>
<td>H₂S</td>
<td>Hydrogen sulfide</td>
<td>Very toxic, odor of rotten eggs</td>
</tr>
<tr>
<td>CO</td>
<td>Carbon monoxide</td>
<td>Toxic, colorless, odorless</td>
</tr>
<tr>
<td>CO₂</td>
<td>Carbon dioxide</td>
<td>Colorless, odorless</td>
</tr>
<tr>
<td>CH₄</td>
<td>Methane</td>
<td>Colorless, odorless, flammable</td>
</tr>
<tr>
<td>C₂H₄</td>
<td>Ethylene</td>
<td>Colorless; ripens fruit</td>
</tr>
<tr>
<td>C₃H₈</td>
<td>Propane</td>
<td>Colorless; bottled gas</td>
</tr>
<tr>
<td>N₂O</td>
<td>Nitrous oxide</td>
<td>Colorless, sweet odor, laughing gas</td>
</tr>
<tr>
<td>NO₂</td>
<td>Nitrogen dioxide</td>
<td>Toxic, red-brown, irritating odor</td>
</tr>
<tr>
<td>NH₃</td>
<td>Ammonia</td>
<td>Colorless, pungent odor</td>
</tr>
<tr>
<td>SO₂</td>
<td>Sulfur dioxide</td>
<td>Colorless, irritating odor</td>
</tr>
</tbody>
</table>
CHARACTERISTICS OF GASES

Gases always form _________ mixtures with other gases.

What does **homogeneous** mean?

Name a **homogeneous** mixture for common gases _________

Hint – what do we breath?
## Composition of Dry Air

<table>
<thead>
<tr>
<th>Name</th>
<th>Formula</th>
</tr>
</thead>
<tbody>
<tr>
<td>Nitrogen</td>
<td>_______</td>
</tr>
<tr>
<td>Oxygen</td>
<td>_______</td>
</tr>
<tr>
<td>Argon</td>
<td>_______</td>
</tr>
<tr>
<td>Carbon Dioxide</td>
<td>_______</td>
</tr>
<tr>
<td>Neon</td>
<td>_______</td>
</tr>
<tr>
<td>Helium</td>
<td>_______</td>
</tr>
<tr>
<td>Krypton</td>
<td>_______</td>
</tr>
<tr>
<td>Xenon</td>
<td>_______</td>
</tr>
</tbody>
</table>
HOW DO WE STUDY GASES?

We “do” ______________

What kind of EXPERIMENTS?
It is common to synthesize gases and collect them by displacing a volume of water.
Which Gases React With Water?

\[
\begin{align*}
N_2 & \quad + \quad H_2O \quad \rightarrow \quad \text{__________} \\
O_2 & \quad + \quad H_2O \quad \rightarrow \quad \text{__________} \\
He & \quad + \quad H_2O \quad \rightarrow \quad \text{__________} \\
Ne & \quad + \quad H_2O \quad \rightarrow \quad \text{__________} \\
Ar & \quad + \quad H_2O \quad \rightarrow \quad \text{__________} \\
Kr & \quad + \quad H_2O \quad \rightarrow \quad \text{__________} \\
Xe & \quad + \quad H_2O \quad \rightarrow \quad \text{__________} \\
CO_2 & \quad + \quad H_2O \quad \rightarrow \quad \text{__________}
\end{align*}
\]
HOW DO WE STUDY GASES?

We ___________ things

WHAT CAN WE MEASURE?

1. PRESSURE of a gas
2. TEMPERATURE of a gas
3. VOLUME of a gas
4. AMOUNT of a gas
Gases exert ________ on the surface they contact.
Definition OF Pressure

FORCE per unit AREA

\[ P = \frac{F}{A} \]

The SI Unit of pressure is the ____________ (Pa)
# Units of Pressure

<table>
<thead>
<tr>
<th>Unit</th>
<th>Conversion Factor</th>
</tr>
</thead>
<tbody>
<tr>
<td>Pascal</td>
<td>$1 \text{ atm} = 101325 \text{ Pa}$</td>
</tr>
<tr>
<td>Torr</td>
<td>$1 \text{ atm} = 760 \text{ Torr}$</td>
</tr>
<tr>
<td>Bar</td>
<td>$1 \text{ atm} = 1.01325 \text{ bar}$</td>
</tr>
<tr>
<td>mmHg</td>
<td>$1 \text{ atm} = 760 \text{ mmHg}$</td>
</tr>
<tr>
<td>lb/in²</td>
<td>$1 \text{ atm} = 14.696 \text{ lb/in}^2$</td>
</tr>
<tr>
<td>in Hg</td>
<td>$1 \text{ atm} = 29.921 \text{ in Hg}$</td>
</tr>
</tbody>
</table>
Atmospheric pressure is measured with a ____________

What Does a BAROMETER Look Like?
Why doesn’t the mercury “run out” of the tube into the dish?
Why doesn’t the mercury “run out” of the tube?

\[
\text{FORCE } 1 = \text{FORCE } 2
\]

\[
\text{FORCE } 1 = \underline{\hspace{2cm}}
\]

\[
\text{FORCE } 2 = \underline{\hspace{2cm}}
\]
Force = mass x acceleration …………………. F = m a

\[
P = \frac{\text{Force}}{\text{Area}} = \frac{(\text{mass})(\text{acceleration})}{\text{Area}}
\]

acceleration (due to gravity) is a constant ……… a = g

\[
\text{Pressure} = \frac{\text{Force}}{\text{Area}} = \frac{(\text{mass})(\text{gravity})}{\text{Area}}
\]
Density = \frac{mass}{volume}

or

mass = density \times volume

\text{Pressure} = \frac{\text{Force}}{\text{Area}} = \frac{(mass)(gravity)}{\text{Area}}

P = \frac{(mass)(g)}{\text{Area}} = \frac{(density)(Volume)(g)}{\text{Area}}
**PRESSURE = FORCE ÷ AREA**

\[ P = \frac{F}{A} = \text{ma} = \frac{mg}{\text{Area}} = \frac{(\text{density})(\text{Volume})(g)}{\text{Area}} \]

but \hspace{1cm} \text{Volume (V) = Area (A) \times height (h)}

\[ P = \frac{(\text{density})(\text{Area})(\text{height})(g)}{\text{Area}} \]

\[ P = g \ d \ h \]
SUMMARY

\[ P = \frac{F}{A} \]

\[ P = \frac{(m \ a)}{A} \]

\[ P = \frac{(d \ V)(g)}{A} \]

\[ P = \frac{(d)(A \ h)(g)}{A} \]

\[ P = g \ d \ h \]
If an oil (whose density is 0.99987 g/cc) was used in the barometer instead of mercury, what would the height of the oil be?

Density of mercury = 13.559 g/cc
Height of mercury = 760 mm
HOW TO WORK PROBLEMS

1. What do you want to know?

2. What do you know?

3. How do you work the problem?
   What “equation” will you use?
The relationship between pressure (P) and the height (h) of a liquid in a barometer is

\[ P = g \cdot d \cdot h \]

where \( g \rightarrow \) constant acceleration of gravity
\( d \rightarrow \) density of liquid in manometer

for mercury: \( P_{\text{Hg}} = \) _____________

For oil: \( P_{\text{Oil}} = \) _____________

\[ P_{\text{Hg}} = P_{\text{Oil}} \]
\[ P = g \frac{d_{\text{oil}} \cdot h_{\text{oil}}}{d_{\text{oil}} \cdot h_{\text{oil}}} = g \frac{d_{\text{Hg}} \cdot h_{\text{Hg}}}{d_{\text{Hg}} \cdot h_{\text{Hg}}} \]

\[ d_{\text{oil}} \cdot h_{\text{oil}} = d_{\text{Hg}} \cdot h_{\text{Hg}} \]

\[ h_{\text{oil}} = d_{\text{Hg}} \cdot h_{\text{Hg}} \div d_{\text{oil}} \]

\[ h_{\text{oil}} = (760.0)(13.559) \div (0.99987) \]

\[ h_{\text{oil}} = \rule{0.5\text{cm}}{0.5\text{cm}} \text{ mm} \]
How do we study gases?

We “do” experiments.

What kind of experiments?

Can measure:

Pressure
Volume
Temperature
Amount
How do you measure the pressure of a gas??

Pressures of gases are measured with ____________

1. OPEN ________________
2. CLOSED ________________
$P_{gas} = P_{atm} + P_{h}$
End Manometers

Measurement dependent on material in manometer

\[ P_{\text{gas}} = P_{\text{atm}} + P_h \]
What is the gas pressure?
If Atmospheric pressure is 1.000 atm & \( h = 67 \text{ mm} \)

\[
P_{\text{gas}} = P_{\text{atm}} + P_{h2}
\]
\[
= 1.000 \text{ atm} + 67 \text{ mm}
\]
\[
= (760.0 + 67) \text{ mm}
\]
\[
= 827 \text{ mm} = 1.088 \text{ atm}
\]

\[
P_{\text{gas}} = \underline{1.088} \text{ atm}
\]
PRESSURE - VOLUME
EXPERIMENTS

BOYLE’S LAW
Results of experiment: When the Pressure doubles the volume is __________
For a **fixed amount** of a gas at a **constant temperature**, pressure is *inversely* proportional to **volume**.
EQUATIONS OF STATE

An Equation Of State Relates The Properties Of A Material To Each Other Mathematically
Pressure is inversely proportional to Volume

\[ P \propto \frac{1}{V} \]

\[ P = (a \text{ constant}) \frac{1}{V} \]

\[ P \ V = A \text{ Constant} \]
BOYLE'S LAW

PRESSURE
is inversely proportional to
VOLUME
at FIXED TEMPERATURE
and AMOUNT

\[ P_1 \ V_1 = P_2 \ V_2 = A \text{  Constant} \]
Volume and Temperature

Charles’ experiments
Charles’ experiments

<table>
<thead>
<tr>
<th>Volume (mL)</th>
<th>Temp (°C)</th>
<th>Temp (K)</th>
</tr>
</thead>
<tbody>
<tr>
<td>30</td>
<td>50</td>
<td>323</td>
</tr>
<tr>
<td>35</td>
<td>100</td>
<td>373</td>
</tr>
</tbody>
</table>
CHARLES’ LAW

VOLUME IS
DIRECTLY PROPORTIONAL

TO ABSOLUTE Temperature
[AT FIXED PRESSURE AND AMOUNT]
**Charles' Law**

\[ V \propto T \]

How do you go from proportional to equal?

\[ V = \text{constant} \times T \]
CHARLES’ LAW

\[ V = \text{constant} \times T \]

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

T MUST be in Kelvin
VOLUME and AMOUNT

Avogadro
### Quantity-Volume Relationship

<table>
<thead>
<tr>
<th></th>
<th>He</th>
<th>N₂</th>
<th>CH₄</th>
</tr>
</thead>
<tbody>
<tr>
<td>Volume</td>
<td>22.4 L</td>
<td>22.4 L</td>
<td>22.4 L</td>
</tr>
<tr>
<td>Pressure</td>
<td>1 atm</td>
<td>1 atm</td>
<td>1 atm</td>
</tr>
<tr>
<td>Temperature</td>
<td>0°C</td>
<td>0°C</td>
<td>0°C</td>
</tr>
<tr>
<td>Mass of gas</td>
<td>4.00 g</td>
<td>28.0 g</td>
<td>16.0 g</td>
</tr>
<tr>
<td>Number 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>
**Avogadro’s Hypothesis:**

Equal volumes of gas

(at the same temperature and pressure)

will contain the same number of molecules
AVOGADRO'S LAW

\[ V_m = 22.4 \, \text{L/mol} \, \text{at STP} \]

Another way of saying that is:

22.4 L of **ANY GAS** at STP contains \(6.02 \times 10^{23}\) gas molecules
What is STP

**STANDARD CONDITIONS**

Standard temperature

0°C = 273.15 K

Standard pressure

1 atm
AVOGADRO’S LAW

\[ V \propto n \]

\[ V = \text{constant} \times n \]

\[ \frac{V_1}{n_1} = \frac{V_2}{n_2} \]
**REVIEW OF GAS LAWS**

- **Boyle’s Law**
  \[ V \propto \frac{1}{P} \] (constant \( n, T \))

- **Charles’s Law**
  \[ V \propto T \] (constant \( n, P \))

- **Avogadro’s Law**
  \[ V \propto n \] (constant \( P, T \))
Review of Gas Laws

Boyle’s Law ............ \( P V = a \) constant

Charles’s Law .......... \( \frac{V}{T} = a \) constant

Avogadro’s Law ........ \( \frac{V}{n} = a \) constant
Combined Gas Laws

\[
\frac{PV}{nT} = a \text{ constant}
\]

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

What is the gas constant \( R \) ?
There is only One gas Equation you need to know

What is it?
THE IDEAL GAS EQUATION

\[ PV = n R T \]
Problem Solving

Pressure–Volume
Pressure 1 = 1.0 atm
Volume 1 = 1.0 Liter

What is the new Volume
If
The Pressure is Doubled
Pressure–Volume

$P = 1.0 \text{ atm}$

$V = 1.0 \text{ L}$

Increase pressure

$P = 2.0 \text{ atm}$

$V = 0.5 \text{ L}$
\[ P_1 V_1 = P_2 V_2 \]

\[ P_1 V_1 = (1.0 \text{ atm})(1.0 \text{ L}) = 1.0 \text{ Lit atm} \]

\[ P_2 V_2 \text{ Must also} = 1.0 \text{ Lit atm} \]

\[ V_2 = \frac{P_1 V_1}{P_2} \]

\[ V_2 = \frac{(1.0 \text{ atm})(1.0 \text{ L})}{(2.0 \text{ atm})} \]

\[ V_2 = \underline{1.0} \text{ Liter} \]
Pressure 1 = 1.0 atm
Volume 1 = 1.0 Liter

What is the new Volume
If
The Pressure is Tripled?
PLOT DATA

1.000 g O₂ at 0°C

Volume (L)

1.40 L

0.70 L

Pressure (atm)

0.25 0.50 0.75 1.00
Problem Solving

Volume -- Temperature
What is the new Volume if The Temperature is Doubled?

<table>
<thead>
<tr>
<th>Temp (°C)</th>
<th>Volume (mL)</th>
</tr>
</thead>
<tbody>
<tr>
<td>50</td>
<td>30</td>
</tr>
<tr>
<td>100</td>
<td>?</td>
</tr>
</tbody>
</table>
**CHARLES' LAW**

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

T MUST be in Kelvin

<table>
<thead>
<tr>
<th>Temp (°C)</th>
<th>Temp (K)</th>
<th>Volume (mL)</th>
</tr>
</thead>
<tbody>
<tr>
<td>50</td>
<td>323</td>
<td>30.0</td>
</tr>
<tr>
<td>100</td>
<td>373</td>
<td>?</td>
</tr>
</tbody>
</table>
\[ V_1 = 30.0 \text{ ml} \quad V_2 = ??? \]
\[ T_1 = 323 \text{ K} \quad T_2 = 373 \text{ K} \]

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

\[ \frac{30.0}{323} = \frac{V_2}{373} \]

\[ V_2 = (373) \frac{30}{323} = 34.6 \]
## Charles’ experiments

<table>
<thead>
<tr>
<th>Volume (mL)</th>
<th>Temperature (°C)</th>
</tr>
</thead>
<tbody>
<tr>
<td>16</td>
<td>-100</td>
</tr>
<tr>
<td>21</td>
<td>- 50</td>
</tr>
<tr>
<td>30</td>
<td>50</td>
</tr>
<tr>
<td>35</td>
<td>100</td>
</tr>
</tbody>
</table>
**Absolute Zero**

- A plot of V versus T is a straight line.
- When T is measured in °C, the intercept on the temperature axis is -273.15°C.
- We define absolute zero, 0 K = -273.15°C.
Problem Solving

Volume -- Amount
Avogadro’s Law: Volume & Amount

The Law of Combining Volumes

For Example

\[ 2 \text{H}_2 (g) + \text{O}_2 (g) \rightarrow 2 \text{H}_2\text{O} (g) \]

Two (2) VOLUMES of Hydrogen react with one (1) VOLUME of Oxygen
THE IDEAL GAS EQUATION

PV = n R T
THE GAS CONSTANT \( R \)

\[
R = \frac{PV}{nT}
\]

At Standard temperature \( 0^\circ C = 273.15 \text{ K} \)
And Standard pressure \( \ldots \ldots \ldots 1 \text{ atm} \)
One mole of any gas occupies \( \ldots 22.4 \text{ Liters} \)
Using these values
THE GAS CONSTANT

\[ PV = nRT \]

\[ R = \frac{PV}{nT} \]

\[ = \frac{(1 \text{ atm})(22.4 \text{ Liters})}{(1.000 \text{ mole})(273.15\text{K})} \]

\[ = 0.08206 \text{ L} \cdot \text{atm/mole} \cdot K \]
<table>
<thead>
<tr>
<th>Units</th>
<th>Numerical Value</th>
</tr>
</thead>
<tbody>
<tr>
<td>L-atm &gt; mol-K</td>
<td>0.08206</td>
</tr>
<tr>
<td>J &gt; mol-K</td>
<td>8.314</td>
</tr>
<tr>
<td>cal &gt; mol-K</td>
<td>1.987</td>
</tr>
<tr>
<td>m³-Pa &gt; mol-K</td>
<td>8.314</td>
</tr>
<tr>
<td>L-torr &gt; mol-K</td>
<td>62.36</td>
</tr>
</tbody>
</table>

\(^a\)SI unit.
The gas equations describe **HOW** gases behave, but they do not explain **WHY** they behave as they do.

The Kinetic-Molecular Theory attempts to explain why.
Kinetic-Molecular Theory

1. Gases consist of molecules that are in constant random motion
2. The volume of all the molecules of the gas is negligible relative to total volume
3. Attractive & repulsive forces are negligible
4. Energy can be transferred between molecules during collision
5. Avg. kinetic energy is proportional to temp
APPLICATIONS OF THE IDEAL GAS EQUATION

1. **GAS DENSITY** ------- **WEIGHT / VOLUME**
2. Molecular Weight ------- grams / mole
3. Partial Pressures
4. Collecting Gases over Water
5. Gas Volumes in Chemical Reactions
Density = weight divided by volume

Where is weight (in grams) found in the Ideal gas Equation

\[ PV = nRT \]

\[ n \text{ (moles)} = \frac{\text{grams}}{\text{grams/mole}} = \frac{\text{grams}}{\text{MW}} \]
**GAS DENSITY**

\[ PV = \frac{\text{grams}}{MW} \quad \text{R} \quad \text{T} \]

\[ P \quad V \quad MW = \text{grams} \quad \text{R} \quad \text{T} \]

\[ \text{Density} = \frac{\text{grams}}{\text{volume}} = \frac{P \quad MW}{R \quad T} \]
Why does Warm Air Rise?
Problem 10.42 Calculate the density of sulfur hexafluoride at 455 torr and 32 °C

1. What do you want to know?
2. What do you know?
3. How do you work the problem?
sulfur hexafluoride at 455 torr and 32 °C

\[
\text{Density} = \frac{\text{grams}}{\text{volume}} = \frac{P}{R} \frac{MW}{T}
\]

Formula \( \text{SF}_6 \)

Molecular Weight = \( 32 + (6)(19) = 146 \)

Pressure (in atm) = \( \frac{455}{760} = 0.5986842 \)

Temp (in Kelvin) = \( 32 + 273 = 305 \)
density at 455 torr and 32 °C

\[ \text{Density} = \frac{\text{grams}}{\text{volume}} = \frac{P}{R} \frac{MW}{T} \]

\[ \text{Density} = \frac{0.5986842}{0.0821} \frac{146}{305} \]

Density = 3.4906606 = ????????????
Density = 3.49 grams/Liter
10.37 Which gas is most dense at STP?

(a) Carbon dioxide
(b) Dinitrogen oxide
(c) Chlorine

Density = \frac{\text{grams}}{\text{volume}} = \frac{P \cdot MW}{R \cdot T}
APPLICATIONS OF THE IDEAL GAS EQUATION

1. Gas Density -------- weight / volume
2. MOLECULAR WEIGHT ---- GRAMS / MOLE
3. Partial Pressures
4. Collecting Gases over Water
5. Gas Volumes in Chemical Reactions
MOLAR MASS (MOLECULAR WEIGHT)

From $P V = n R T$

$$n = \text{moles} = \frac{\text{grams}}{MW}$$

$$MW = \frac{gRT}{VP}$$
Problem 10.50 An unknown vapor had a mass of 0.846 grams a volume of 354 cm$^3$, at 752 torr when temp was 100 °C.

CALCULATE THE MOLAR MASS OF THE UNKNOWN VAPOR.
HOW TO WORK PROBLEMS

1. What do you want to know ?
   Molecular Weight

2. What do you know ?
   Pressure, Temperature, Volume and amount

3. How do you work the problem?
   IDEAL GAS LAW
MOLECULAR WEIGHT OF A GAS

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

\[ n = \text{moles} = \frac{\text{grams}}{MW} \]

\[ MW = \frac{gRT}{VP} \]
grams = 0.846 ; volume = 354 cm³; pressure = 752 torr; temperature = 100 °C.

\[
MW = \frac{gRT}{VP} \\
MW = \frac{(0.846)(0.0821)(100 + 273.15)}{(0.354)(752 / 760)}
\]

MW = 73.9 grams / mol
16 grams of an unknown diatomic gas occupies 11.2 liters at STP. What is it?

1. What do you want to know?  
   Molecular Weight

2. What do you know?  
   Pressure, temperature & Volume

3. How do you work the problem?  
   \[ PV = n \cdot RT \quad \text{and} \quad MW = \frac{\text{grams}}{\text{moles}} \]
16 grams of an unknown diatomic gas occupies 11.2 liters at STP. What is it?

Using \[ MW = \frac{\text{grams}}{\text{moles}} \]

You already know grams, so only need moles.

From \[ PV = nRT \] calculate moles:

What is the easiest way to do that?

\[ MW = \frac{\text{grams}}{\text{moles}} = 16 / \frac{1}{2} = 32 \]

Gas is probably oxygen.
APPLICATIONS OF THE IDEAL GAS EQUATION

1. Gas Density -------------- weight / volume
2. Molecular Weight -------- grams / mole

3. PARTIAL PRESSURES

4. Collecting Gases over Water
5. Gas Volumes in Chemical Reactions
Dalton’s Law: in a gas mixture the total pressure is given by the sum of partial pressures of each component:

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

\[ P_i = n_i \left( \frac{RT}{V} \right) \]
Gas Mixtures and Partial Pressures

\[ P_{\text{total}} = (n_1 + n_2 + n_3 + \cdots) \left( \frac{RT}{V} \right) \]

Taking a Ratio

\[ \frac{P_i}{P_T} \frac{V}{V} = \frac{n_i}{n_T} \frac{R}{R} \frac{T}{T} \]

\[ \frac{P_i}{P_T} = \frac{n_i}{n_T} \]
Gas Mixtures and Partial Pressures

\[
\frac{P_i}{P_T} = \frac{n_i}{n_T} \tag{1}
\]

\[
P_i = X_i P_{\text{total}}
\]

where \(X_i\) is the **mole fraction**

\[
X_i = \frac{n_i}{n_{\text{Total}}}
\]
A gas mixture contains 0.80 mol N\textsubscript{2} and 0.20 mol O\textsubscript{2}.

What is the partial pressure of N\textsubscript{2} if the total pressure is 1.0 atm?

\[ P_{\text{total}} = P_{N_2} + P_{O_2} \]

\[ \frac{P_{N_2}}{P_T} \cdot \frac{V}{V} = \frac{n_{N_2}}{n_T} \cdot \frac{R}{R} \cdot \frac{T}{T} \]
0.80 mol N₂, 0.20 mol O₂ and P_T 1.0 atm

\[
\frac{P_{N_2}}{P_T} \frac{V}{V} = \frac{n_{N_2}}{n_T} \frac{RT}{RT}
\]

\[
\frac{P_{N_2}}{P_T} = \frac{n_{N_2}}{n_T} = X_{N_2}
\]

\[
P_{N_2} = X_{N_2} P_T = \frac{0.80}{0.80 + 0.20} (1.0 \text{ atm}) = 0.80 \text{ atm}
\]
A gas mixture contains 22.4 gms $N_2$ and 6.4 gms $O_2$. What is the partial pressure of $N_2$ if the total pressure is 1.0 atm.

SAME problem!

22.4 grams nitrogen = 0.80 mol
6.4 grams oxygen = 0.20 mol
# Vapor Pressure of Water

<table>
<thead>
<tr>
<th>$T , (^\circ \text{C})$</th>
<th>$P , (\text{torr})$</th>
<th>You are NOT required to calculate or know the vapor pressure of water EXCEPT at 100 $^\circ \text{C}$ and one atm.</th>
</tr>
</thead>
<tbody>
<tr>
<td>0</td>
<td>4.58</td>
<td></td>
</tr>
<tr>
<td>10</td>
<td>9.21</td>
<td></td>
</tr>
<tr>
<td>20</td>
<td>17.54</td>
<td></td>
</tr>
<tr>
<td>25</td>
<td>23.76</td>
<td></td>
</tr>
<tr>
<td>30</td>
<td>31.82</td>
<td></td>
</tr>
<tr>
<td>50</td>
<td>92.5</td>
<td></td>
</tr>
<tr>
<td>100</td>
<td>??????</td>
<td></td>
</tr>
</tbody>
</table>
How many moles of oxygen would be obtained from 1.00 mole of Potassium chlorate?

2 \( \text{KClO}_3(\text{s}) \rightarrow 3 \text{O}_2(\text{gas}) + 2 \text{KCl(} \text{s}) \)

or

1 \( \text{KClO}_3 \rightarrow 1 \frac{1}{2} \text{O}_2 + 1 \text{KCl} \)
If this amount of oxygen were collected at STP, what volume would it occupy?

\[ P \ V = n \ R \ T \ \Rightarrow \ V = \frac{n \ o_2 \ RT}{P \ o_2} \]

\[ V = \frac{1\frac{1}{2} \ (0.0821)(0 + 273)}{P \ o_2} \]
How do you change torr to atm?

\[ P_{gas} = P_{total} - P_{water} \]

\[ P_{O_2} = 1.00 \text{ atm} - 4.58 \text{ torr} \]

\[ V = \frac{1\frac{1}{2} (0.0821)(0+273)}{P_{O_2}} \]
APPLICATIONS OF THE IDEAL GAS EQUATION

1. Gas Density ---------------- weight / volume
2. Molecular Weight --------- grams / mole
3. Partial Pressures
4. GAS VOLUMES IN CHEMICAL REACTIONS
What volume of ammonia would be formed if 1.2 L of N\textsubscript{2} reacts completely with 3.6 L of H\textsubscript{2}?

\textit{1st Write and Balance Reaction}

\[ \text{N}_2 + 3 \text{H}_2 \rightarrow 2 \text{NH}_3 \]

\textit{Next, Interpret Equation}

One mole (or one volume) of N\textsubscript{2} Reacts with 3 moles (or 3 volumes) of H\textsubscript{2} To Produce 2 moles (or 2 volumes) of NH\textsubscript{3}

Or in this problem, \boxed{2.4 \text{ Liters}} of NH\textsubscript{3}
Further Applications

Putting Concepts Together
COLLECTING GASES OVER WATER

Does the gas REACT with water?

Which of the following react with water?

Oxygen + water →

Carbon diOxide + water →

Methane + water →

Ammonia + water →
COLLECTING GASES OVER WATER

- It is common to synthesize gases and collect them by displacing a volume of
Heating Potassium chlorate produces Oxygen and Potassium chloride.

1st Write Reaction

\[ \text{KClO}_3(s) + \text{heat} \rightarrow \text{O}_2 \text{(gas)} + \text{KCl} \text{(s)} \]

2nd Balance Reaction

\[ 2 \text{ KClO}_3(s) \rightarrow 3 \text{ O}_2 \text{(gas)} + 2 \text{ KCl} \text{(s)} \]
Water has a vapor pressure which **varies with temperature**. To determine the pressure of a gas collected over water must correct for the partial pressure of the water:

\[
P_{\text{total}} = P_{\text{gas}} + P_{\text{water}}
\]

\[
P_{\text{gas}} = P_{\text{total}} - P_{\text{water}}
\]
What volume (at STP) of O₂ would be obtained from 1.00 mole of KClO₃

$$1 \text{ KClO}_3 \rightarrow 1 \frac{1}{2} \text{ O}_2 + 1 \text{ KCl}$$

$$1 \frac{1}{2} = 22.4 + 11.2 = 33.6 \text{ Liters}$$
The oxidation of glucose in our bodies produces CO$_2$ which is expelled from our lungs

- Calculate the volume of CO$_2$ produced at body temperature (37 °C) and 0.970 atm when 25.5 g of C$_6$H$_{12}$O$_6$ is consumed.
- What do you do first?
- Right! Write & Balance reaction
\[
C_6H_{12}O_6 + O_2 \rightarrow \text{??????}
\]

\[
C_6H_{12}O_6 + O_2 \rightarrow \text{CO}_2(\text{gas}) + \text{H}_2\text{O(liq)}
\]
Balanced ???????

\[
1 \ C_6H_{12}O_6 + 6 \ O_2 \rightarrow 6 \ \text{CO}_2(\text{gas}) + 6 \ \text{H}_2\text{O(liq)}
\]
Now what?

Right, calculate moles of \(C_6H_{12}O_6\)

\[
24.5 \ \text{grams} \times \text{???} = \text{moles}
\]
\[ 1 \text{ C}_6\text{H}_{12}\text{O}_6 + 6 \text{ O}_2 \rightarrow 6 \text{ CO}_2(\text{gas}) + 6 \text{ H}_2\text{O}(\text{liq}) \]

\[
\frac{24.5 \text{ g}}{180 \text{ g/mol}} = 0.136 \text{ mol glucose}
\]

\[
(0.136) \times 6 = \text{mole CO}_2 \text{ produced} = 0.816
\]

From \( P \ V = n \ R \ T \)

\[
V = \frac{n \ R \ T}{P}
\]

\[
V = (0.816)(0.08206)(37 + 273) / 0.970
\]

\[
V = 21.4 \text{ Liters CO}_2 \text{ produced}
\]
Effusion

The escape of gas molecules through a tiny hole into an evacuated space.
**GRAHAM’S LAW OF EFFUSION**

\[
\frac{r_1}{r_2} = \sqrt{\frac{MW_2}{MW_1}}
\]

Gas Separations (page 427 text)

- The slight difference in molar mass between $^{235}\text{UF}_6$ and $^{238}\text{UF}_6$ causes the molecules to move at different rates:

\[
\frac{r_{235}}{r_{238}} = \sqrt{\frac{MW_{238}}{MW_{235}}} = 1.0043
\]
Chapter 10 (gases)

The rest of the story
Real Gases

Deviations from Ideal Behavior
For an ideal gas

1. the molecules of a gas have NO volume

2. molecules of a gas do NOT attract (or repel) each other.
Real Gases: Deviations from Ideal Behavior

(a)  

(b)
Pressure

As the pressure on a gas increases, the molecules are forced closer together.

As the gas molecules get closer together the more likely attractive forces will develop between the molecules.
As the molecules get closer together, the volume of the container gets smaller.

The smaller the container, the more space the gas molecules occupy.
Temperature

• As temperature increases, the gas molecules move faster and further apart.

• Also, higher temperatures mean more energy available to break intermolecular forces.
Real Gases:
Deviations from Ideal Behavior

The higher the pressure, the more NON ideal

Why?

The higher the pressure the more gas molecules therefore the more interaction between molecules and more volume the gas occupies
Real Gases: Deviations from Ideal Behavior

The smaller the volume, the more NON ideal

Why?

Molecules are closer together and therefore more interaction between molecules and more volume the gas occupies
Real Gases: Deviations from Ideal Behavior

The higher the temperature, the more ideal the gas.

Why?

Because the gas molecules have less interaction due to thermal motion.
The van der Waals Equation

- Two terms are added to the ideal gas equation
- one to correct for *volume* of molecules
- and the other to correct for *(pressure)* intermolecular attractions
The van der Waals Equation

\[ P = \frac{nRT}{V - nb} - \frac{n^2a}{V^2} \]

Corrects for molecular volume

Corrects for molecular attraction
End of Chapter 10: Gases