Chapter 9

The Mole Concept

by Christopher Hamaker
Avogadro’s Number

• *Avogadro’s number* (symbol \( N \)) is the number of atoms in 12.01 grams of carbon.

• Its numerical value is \( 6.02 \times 10^{23} \).

• Therefore, a 12.01 g sample of carbon contains \( 6.02 \times 10^{23} \) carbon atoms.
The Mole

- The mole (mol) is a unit of measure for an amount of a chemical substance.

- A mole is Avogadro’s number of particles, which is $6.02 \times 10^{23}$ particles.

\[ 1 \text{ mol} = \text{Avogadro’s number} = 6.02 \times 10^{23} \text{ units} \]

- We can use the mole relationship to convert between the number of particles and the mass of a substance.
Analogies for Avogadro’s Number

- The volume occupied by one mole of softballs would be about the size of Earth.
- One mole of Olympic shot put balls has about the same mass as that of Earth.
- One mole of hydrogen atoms laid side by side would circle Earth about 1 million times.
One Mole of Several Substances

- C\textsubscript{12}H\textsubscript{22}O\textsubscript{11}
- H\textsubscript{2}O
- mercury
- sulfur
- NaCl
- K\textsubscript{2}Cr\textsubscript{2}O\textsubscript{7}
- lead
- copper
Mole Calculations

• We will be using the unit analysis method again.

• Recall the following steps:

1. Write down the unit requested.

2. Write down the given value.

3. Apply unit factor(s) to convert the given units to the desired units.
Mole Calculations 1

- How many sodium atoms are in 0.120 mol Na?
  1. We want atoms of Na.
  2. We have 0.120 mol Na.
  3. 1 mole Na = 6.02 \times 10^{23} \text{ atoms Na}.

\[
0.120 \text{ mol Na} \times \frac{6.02 \times 10^{23} \text{ atoms Na}}{1 \text{ mol Na}} = 7.22 \times 10^{22} \text{ atoms Na}
\]
• How many moles of potassium are in $1.25 \times 10^{21}$ atoms K?

1. We want moles K.

2. We have $1.25 \times 10^{21}$ atoms K.

3. 1 mole K = $6.02 \times 10^{23}$ atoms K.

\[
1.25 \times 10^{21} \text{ atoms K} \times \frac{1 \text{ mol K}}{6.02 \times 10^{23} \text{ atoms K}} = 2.08 \times 10^{-3} \text{ mol K}
\]
Molar Mass

• The atomic mass of any substance expressed in grams is the *molar mass* (MM) of that substance.

• The atomic mass of iron is 55.85 amu.

• Therefore, the molar mass of iron is 55.85 g/mol.

• Since oxygen occurs naturally as a diatomic, O₂, the molar mass of oxygen gas is two times 16.00 g or 32.00 g/mol.
Calculating Molar Mass

• The molar mass of a substance is the sum of the molar masses of each element.

• What is the molar mass of magnesium nitrate, Mg(NO$_3$)$_2$?

• The sum of the atomic masses is as follows:

\[
24.31 + 2(14.01 + 16.00 + 16.00 + 16.00) =
24.31 + 2(62.01) = 148.33 \text{ amu}
\]

• The molar mass for Mg(NO$_3$)$_2$ is 148.33 g/mol.
• Now we will use the molar mass of a compound to convert between grams of a substance and moles or particles of a substance.

\[ 6.02 \times 10^{23} \text{ particles} = 1 \text{ mol} = \text{molar mass} \]

• If we want to convert particles to mass, we must first convert particles to moles, and then we can convert moles to mass.
What is the mass of 1.33 moles of titanium, Ti?

1. We want grams.

2. We have 1.33 moles of titanium.

3. Use the molar mass of Ti: 1 mol Ti = 47.88 g Ti.

\[
1.33 \text{ mole Ti} \times \frac{47.88 \text{ g Ti}}{1 \text{ mole Ti}} = 63.7 \text{ g Ti}
\]
Mole Calculations III

• What is the mass of $2.55 \times 10^{23}$ atoms of lead?

1. We want grams.
2. We have atoms of lead.
3. Use Avogadro’s number and the molar mass of Pb.

\[
2.55 \times 10^{23} \text{ atoms Pb} \times \frac{1 \text{ mol Pb}}{6.02 \times 10^{23} \text{ atoms Pb}} \times \frac{207.2 \text{ g Pb}}{1 \text{ mole Pb}} = 87.9 \text{ g Pb}
\]
How many \( \text{O}_2 \) molecules are present in 0.470 g of oxygen gas?

1. We want molecules \( \text{O}_2 \).
2. We have grams \( \text{O}_2 \).
3. Use Avogadro’s number and the molar mass of \( \text{O}_2 \).

\[
0.470 \text{ g} \text{O}_2 \times \frac{1 \text{ mol} \text{O}_2}{32.00 \text{ g} \text{O}_2} \times \frac{6.02 \times 10^{23} \text{ molecules} \text{O}_2}{1 \text{ mole} \text{O}_2} = 8.84 \times 10^{21} \text{ molecules} \text{O}_2
\]
Mass of a Single Molecule

• What is the mass of a single molecule of sulfur dioxide? The molar mass of SO$_2$ is 64.07 g/mol.

• We want mass/molecule SO$_2$, we have the molar mass of sulfur dioxide.

• Use Avogadro’s number and the molar mass of SO$_2$ as follows:

\[
\frac{64.07 \text{ g SO}_2}{1 \text{ mol SO}_2} \times \frac{1 \text{ mole SO}_2}{6.02 \times 10^{23} \text{ molecules SO}_2} = 1.06 \times 10^{-22} \text{ g/molecule}
\]
Molar Volume

- At standard temperature and pressure, 1 mole of any gas occupies 22.4 L.
- The volume occupied by 1 mole of gas (22.4 L) is called the molar volume.
- Standard temperature and pressure are 0 °C and 1 atm.
Molar Volume of Gases

- We now have a new unit factor equation:

\[ 1 \text{ mole gas} = 6.02 \times 10^{23} \text{ molecules gas} = 22.4 \text{ L gas} \]

<table>
<thead>
<tr>
<th>GAS</th>
<th>NO. OF MOLES</th>
<th>NO. OF MOLECULES</th>
<th>MOLAR MASS</th>
<th>MOLAR VOLUME AT STP</th>
</tr>
</thead>
<tbody>
<tr>
<td>hydrogen, H(_2)</td>
<td>1.00</td>
<td>(6.02 \times 10^{23})</td>
<td>2.02 g/mol</td>
<td>22.4 L/mol</td>
</tr>
<tr>
<td>oxygen, O(_2)</td>
<td>1.00</td>
<td>(6.02 \times 10^{23})</td>
<td>32.00 g/mol</td>
<td>22.4 L/mol</td>
</tr>
<tr>
<td>carbon dioxide, CO(_2)</td>
<td>1.00</td>
<td>(6.02 \times 10^{23})</td>
<td>44.01 g/mol</td>
<td>22.4 L/mol</td>
</tr>
<tr>
<td>ammonia, NH(_3)</td>
<td>1.00</td>
<td>(6.02 \times 10^{23})</td>
<td>17.04 g/mol</td>
<td>22.4 L/mol</td>
</tr>
<tr>
<td>argon, Ar(*)</td>
<td>1.00</td>
<td>(6.02 \times 10^{23})</td>
<td>39.95 g/mol</td>
<td>22.4 L/mol</td>
</tr>
</tbody>
</table>

*Argon gas is composed of atoms rather than molecules.

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One Mole of a Gas at STP

- The box below has a volume of 22.4 L, which is the volume occupied by 1 mole of a gas at STP.
Gas Density

• The density of gases is much less than that of liquids.

• We can easily calculate the density of any gas at STP.

• The formula for gas density at STP is as follows:

\[
\frac{\text{molar mass in grams}}{\text{molar volume in liters}} = \text{density, g/L}
\]
Calculating Gas Density

- What is the density of ammonia gas, NH₃, at STP?
- First we need the molar mass for ammonia.
  \[ 14.01 + 3(1.01) = 17.04 \text{ g/mol} \]
- The molar volume NH₃ at STP is 22.4 L/mol.
- Density is mass/volume.
  \[ \frac{17.04 \text{ g/mol}}{22.4 \text{ L/mol}} = 0.761 \text{ g/L} \]
Molar Mass of a Gas

- We can also use molar volume to calculate the molar mass of an unknown gas.

- 1.96 g of an unknown gas occupies 1.00 L at STP. What is the molar mass?

- We want g/mol; we have g/L.

\[
\frac{1.96 \text{ g}}{1.00 \text{ L}} \times \frac{22.4 \text{ L}}{1 \text{ mole}} = 43.9 \text{ g/mol}
\]
Mole Unit Factors

• We now have three interpretations for the mole:
  1. 1 mol = 6.02 x 10^{23} particles.
  2. 1 mol = molar mass.
  3. 1 mol = 22.4 L at STP for a gas.

• This gives us three unit factors to use to convert among moles, particles, mass, and volume.
Calculating Molar Volume

• A sample of methane, CH$_4$, occupies 4.50 L at STP. How many moles of methane are present?

• We want moles; we have volume.

• Use molar volume of a gas: 1 mol = 22.4 L.

$$\frac{4.50 \text{ L CH}_4}{22.4 \text{ L CH}_4} \times \frac{1 \text{ mol CH}_4}{22.4 \text{ L CH}_4} = 0.201 \text{ mol CH}_4$$
Calculating Mass Volume

- What is the mass of 3.36 L of ozone gas, $O_3$, at STP?
- We want mass $O_3$; we have 3.36 L $O_3$.
- Convert volume to moles, then moles to mass.

\[
3.36 \text{ L } O_3 \times \frac{1 \text{ mol } O_3}{22.4 \text{ L } O_3} \times \frac{48.00 \text{ g } O_3}{1 \text{ mol } O_3} = 7.20 \text{ g } O_3
\]
Calculating Molecule Volume

• How many molecules of hydrogen gas, $H_2$, occupy 0.500 L at STP?

• We want molecules $H_2$; we have 0.500 L $H_2$.

• Convert volume to moles, and then moles to molecules.

$$0.500 \text{ L} \ H_2 \times \frac{1 \text{ mol} \ H_2}{22.4 \text{ L} \ H_2} \times \frac{6.02 \times 10^{23} \text{ molecules} \ H_2}{1 \text{ mole} \ H_2}$$

$$= 1.34 \times 10^{22} \text{ molecules} \ H_2$$
Percent Composition

• The *percent composition* of a compound lists the mass percent of each element.

• For example, the percent composition of water, $\text{H}_2\text{O}$ is 11% hydrogen and 89% oxygen.

• All water contains 11% hydrogen and 89% oxygen by mass.
Calculating Percent Composition

• There are a few steps to calculating the percent composition of a compound. Let’s practice using H$_2$O.

1. Assume you have 1 mole of the compound.

2. One mole of H$_2$O contains 2 mol of hydrogen and 1 mol of oxygen. Therefore,

\[2(1.01 \text{ g H}) + 1(16.00 \text{ g O}) = \text{molar mass H}_2\text{O}\]

\[2.02 \text{ g H} + 16.00 \text{ g O} = 18.02 \text{ g H}_2\text{O}\]
Next, find the percent composition of water by comparing the masses of hydrogen and oxygen in water to the molar mass of water.

\[
\frac{2.02 \text{ g H}}{18.02 \text{ g H}_2\text{O}} \times 100\% = 11.2\% \text{ H}
\]

\[
\frac{16.00 \text{ g O}}{18.02 \text{ g H}_2\text{O}} \times 100\% = 88.79\% \text{ O}
\]
Percent Composition Problem

• TNT (trinitrotoluene) is a white crystalline substance that explodes at 240 °C. Calculate the percent composition of TNT, C₇H₅(NO₂)₃.

\[ 7(12.01 \text{ g C}) + 5(1.01 \text{ g H}) + 3 (14.01 \text{ g N} + 32.00 \text{ g O}) = \text{g} \ C₇H₅(NO₂)₃ \]

\[ 84.07 \text{ g C} + 5.05 \text{ g H} + 42.03 \text{ g N} + 96.00 \text{ g O} = 227.15 \text{ g} \ C₇H₅(NO₂)₃ \]
Percent Composition of TNT

\[
\frac{84.07 \text{ g C}}{227.15 \text{ g TNT}} \times 100\% = 37.01\% \text{ C}
\]

\[
\frac{1.01 \text{ g H}}{227.15 \text{ g TNT}} \times 100\% = 2.22\% \text{ H}
\]

\[
\frac{42.03 \text{ g N}}{227.15 \text{ g TNT}} \times 100\% = 18.50\% \text{ N}
\]

\[
\frac{96.00 \text{ g O}}{227.15 \text{ g TNT}} \times 100\% = 42.26\% \text{ O}
\]
Empirical Formulas

• The *empirical formula* of a compound is the simplest whole number ratio of ions in a formula unit or atoms of each element in a molecule.

• The molecular formula of benzene is $C_6H_6$.
  – The empirical formula of benzene is $CH$.

• The molecular formula of octane is $C_8H_{18}$.
  – The empirical formula of octane is $C_4H_9$. 
Calculating Empirical Formulas

• We can calculate the empirical formula of a compound from its composition data.

• We can determine the mole ratio of each element from the mass to determine the empirical formula of radium oxide, Ra₂O₃.
  – A 1.640 g sample of radium metal was heated to produce 1.755 g of radium oxide. What is the empirical formula?
  – We have 1.640 g Ra and 1.755-1.640 = 0.115 g O.
– The molar mass of radium is 226.03 g/mol, and the molar mass of oxygen is 16.00 g/mol.

\[
\begin{align*}
1.640 \text{ g Ra} & \times \frac{1 \text{ mol Ra}}{226.03 \text{ g Ra}} = 0.00726 \text{ mol Ra} \\
0.115 \text{ g O} & \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 0.00719 \text{ mol O}
\end{align*}
\]

– We get Ra$_{0.00726}$O$_{0.00719}$. Simplify the mole ratio by dividing by the smallest number.

– We get Ra$_{1.01}$O$_{1.00}$ = RaO is the empirical formula.
Empirical Formulas from Percent Composition

• We can also use percent composition data to calculate empirical formulas.

• Assume that you have 100 grams of sample.

• Acetylene is 92.2% carbon and 7.83% hydrogen. What is the empirical formula?
  – If we assume 100 grams of sample, we have 92.2 g carbon and 7.83 g hydrogen.
Empirical Formula for Acetylene

• Calculate the moles of each element.

\[
92.2 \text{ g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} = 7.68 \text{ mol C}
\]

\[
7.83 \text{ g H} \times \frac{1 \text{ mol H}}{1.01 \text{ g H}} = 7.75 \text{ mol H}
\]

• The ratio of elements in acetylene is \(C_{7.68}H_{7.75}\). Divide by the smallest number to get the following formula:

\[
C_{\frac{7.68}{7.68}}H_{\frac{7.75}{7.68}} = C_{1.00}H_{1.01} = CH
\]
• The empirical formula for acetylene is CH. This represents the ratio of C to H atoms on acetylene.

• The actual *molecular formula* is some multiple of the empirical formula, \((\text{CH})_n\).

• Acetylene has a molar mass of 26 g/mol. Find \(n\) to find the molecular formula:

\[
\frac{(\text{CH})_n}{\text{CH}} = \frac{26 \text{ g/mol}}{13 \text{ g/mol}} \quad n = 2 \quad \text{and the molecular formula is } C_2H_2.
\]
Chapter Summary

• *Avogadro’s number* is $6.02 \times 10^{23}$, and is 1 mole of any substance.

• The *molar mass* of a substance is the sum of the atomic masses of each element in the formula.

• At STP, 1 mole of any gas occupies 22.4 L.
• We can use the following flow chart for mole calculations:

(a) Apply *Avogadro’s number* as a unit factor: 1 mol = $6.02 \times 10^{23}$ particles.
(b) Apply *molar mass* as a unit factor: 1 mol = ? g (refer to Periodic Table).
(c) Apply *molar volume* as a unit factor: 1 mol = 22.4 L at STP.
• The *percent composition* of a substance is the mass percent of each element in that substance.

• The *empirical formula* of a substance is the simplest whole number ratio of the elements in the formula.

• The *molecular formula* is a multiple of the empirical formula.