Chemistry I: Using Chemical Formulas

**Formula Mass** – *The sum of the average atomic masses of all elements in the compound. Units are amu.*

**Molar Mass** - *The mass in grams of 1 mole of a substance. Substance can be atoms, molecules, formula units, or ions. Units are g/mol.*

**Molecular Mass** – *The mass in grams of 1 mole of a molecular substance. Units are g/mol. Only used for molecules.*

**Molecular Weight** – *Old fashioned term that just won’t die. Means the same as Molar Mass. Units are g/mol.*

Calculate the Molar Mass of the following compounds:

- \( \text{H}_2\text{O} \)  \( 18.02 \text{ g/mol} \)
- \( \text{NaCl} \)  \( 58.44 \text{ g/mol} \)
- \( \text{C}_6\text{H}_{12}\text{O}_6 \)  \( 180.18 \text{ g/mol} \)
- \( \text{HC}_2\text{H}_3\text{O}_2 \)  \( 60.06 \text{ g/mol} \)
- \( \text{CuSO}_4 \cdot 5\text{H}_2\text{O} \)  \( 249.72 \text{ g/mol} \)
Molar Mass can be used as a conversion factor to determine the number of moles of a substance from a given mass, or the mass of a substance from a given number of moles.

**To find Mass in grams from moles:**

**USE THE GRID TO CANCEL UNITS!!!!**

<table>
<thead>
<tr>
<th># of moles</th>
<th>(mass on periodic table) grams</th>
<th>=</th>
<th>_____ mass in grams</th>
</tr>
</thead>
<tbody>
<tr>
<td>1 mole</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

Example 1. What is the mass of 5.00 moles of Nitrogen gas?

\[
5.00 \text{ mol } N_2 \times \frac{28.02 \text{ g } N_2}{1 \text{ mol } N_2} = 1.40 \times 10^2 \text{ g } N_2
\]

**To find Moles from mass in grams:**

<table>
<thead>
<tr>
<th># of grams</th>
<th>1 mole</th>
<th>=</th>
<th>_____ moles</th>
</tr>
</thead>
<tbody>
<tr>
<td>(mass on periodic table) grams</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

Example 2. How many moles of Ammonium Dichromate are contained in 5.50 grams of the substance?

\[
5.50 \text{ g } (\text{NH}_4)_2 \text{Cr}_2\text{O}_7 \times \frac{1 \text{ mol } (\text{NH}_4)_2 \text{Cr}_2\text{O}_7}{252.10 \text{ g } (\text{NH}_4)_2 \text{Cr}_2\text{O}_7} = 0.218 \text{ mol } (\text{NH}_4)_2 \text{Cr}_2\text{O}_7
\]
Counting Particles in Compounds
Consider a 0.750 mole sample of CH$_4$

Calculate the mass, moles, and molecules of the compound, and the mass, moles, and atoms of each element in a given amount of a compound.

1. Moles of Compound → Mass of Compound:

Find the Molar Mass: 1(12.01) + 4(1.10) g/mol = 16.41 g/mol CH$_4$

\[
\frac{0.750 \text{ mole CH}_4}{1 \text{ mole CH}_4} = \frac{16.41 \text{ gCH}_4}{1 \text{ mole}} = 12.3 \text{ g CH}_4
\]

2. Moles of Compound → Moles of Elements in the compound:

*Carbon:*

\[
\frac{0.750 \text{ mole CH}_4}{1 \text{ mole CH}_4} \times \frac{1 \text{ mole C}}{1 \text{ mole CH}_4} = 0.750 \text{ mole C}
\]

*Hydrogen:*

\[
\frac{0.750 \text{ mole CH}_4}{1 \text{ mole CH}_4} \times \frac{4 \text{ mole H}}{1 \text{ mole CH}_4} = 3.00 \text{ mole H}
\]

3. Moles of Element in the Compound → Mass of Elements in compound:

Use Molar Mass of the Element:

*Grams Carbon:*

\[
\frac{0.750 \text{ mole C}}{1 \text{ mole C}} \times \frac{12.01 \text{ g C}}{1 \text{ mole C}} = 9.01 \text{ g C}
\]

Grams Hydrogen:

\[
\frac{3.00 \text{ mole H}}{1 \text{ mole H}} \times \frac{1.01 \text{ g H}}{1 \text{ mole H}} = 3.03 \text{ g Hydrogen}
\]

Notice that the mass of the individual elements must add up to the mass of the compound. (It’s close, but off .01 gram because of sig fig rounding...)
4. Moles of each element in compound → Atoms of each element in compound:

Use Avogadro’s Number: $6.022 \times 10^{23}$ atoms = 1 mole of atoms

**Atoms Carbon:**

$$\frac{0.750 \text{ mole } C}{1 \text{ mole } C} \times \frac{6.022 \times 10^{23} \text{ atoms}}{1 \text{ mole } C} = 4.52 \times 10^{23} \text{ atoms Carbon}$$

**Atoms Hydrogen:**

$$\frac{3.00 \text{ mole } H}{1 \text{ mole } H} \times \frac{6.022 \times 10^{23} \text{ atoms}}{1 \text{ mole } H} = 1.81 \times 10^{24} \text{ atoms Hydrogen}$$

5. Moles of Molecules of the compound → Molecules of the compound

Use Avogadro’s Number: $6.022 \times 10^{23}$ molecules = 1 mole of molecules

$$\frac{0.750 \text{ mole } CH_4}{1 \text{ mole } CH_4} \times \frac{6.022 \times 10^{23} \text{ molecules}}{1 \text{ mole } CH_4} = 4.52 \times 10^{23} \text{ CH}_4 \text{ molecules}$$

### Mass Mole Atom Conversion Practice

Fill in the following table

<table>
<thead>
<tr>
<th>1</th>
<th>2</th>
<th>3</th>
<th>4</th>
<th>5</th>
</tr>
</thead>
<tbody>
<tr>
<td>Grams of Compound</td>
<td>Moles of each element</td>
<td>Grams of each element</td>
<td>Atoms of each element</td>
<td>Molecules of compound</td>
</tr>
<tr>
<td>1.20 moles H₂O MM = 18.02 g/mol</td>
<td>21.6 g</td>
<td>2.40 mole H 1.20 mole O</td>
<td>2.42 g H 19.2 gO</td>
<td>1.44 x 10^{24} 7.22 x 10^{23}</td>
</tr>
<tr>
<td>1.75 moles H₂SO₄ MM = 98.09 g/mol</td>
<td>171.5g</td>
<td>3.50 mol H 1.75 mol S 7.00 mol O</td>
<td>3.54gH 56.1gS 112g O</td>
<td>2.11x10^{24} H 1.05x10^{24} S 4.21x10^{24} O</td>
</tr>
<tr>
<td>( \text{moles} ) ( \text{N}_2\text{O}_4 )</td>
<td>( \text{MM} = 92.02 \text{ g/mol} )</td>
<td>( 5.00 )</td>
<td>( 460 \text{ g} )</td>
<td>( 10.0 \text{ mol N} )</td>
</tr>
<tr>
<td>---</td>
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<td>---</td>
<td>---</td>
</tr>
<tr>
<td>( \text{moles} ) ( \text{BCl}_3 )</td>
<td>( \text{MM} = 117.1 \text{ g/mol} )</td>
<td>( 0.625 )</td>
<td>( 73.2 \text{ g} )</td>
<td>( 0.625 \text{ mol B} )</td>
</tr>
<tr>
<td>( \text{moles} ) ( \text{C}<em>{12}\text{H}</em>{22}\text{O}_{11} )</td>
<td>( \text{MM} = 342.34 \text{ g/mol} )</td>
<td>( 1.50 )</td>
<td>( 513.51 \text{ g} )</td>
<td>( 18.0 \text{ mol C} )</td>
</tr>
</tbody>
</table>

**Finding % Composition of a Formula** *(units= g element/100 g compound)*

**Example:** Find the % Composition of \( \text{Fe}_2\text{O}_3 \)

1. **Start with 1 mole of the compound:**

   *In 1 mole \( \text{Fe}_2\text{O}_3 \) (formula units) there are 2 mole Fe (atoms) and 3 mole O (atoms)*

2. **Determine the moles of each element in the compound:**

   \[
   \begin{align*}
   2 \text{ mole Fe} \\
   3 \text{ mole O}
   \end{align*}
   \]

3. **Convert the moles of each element to grams:**

   \[
   \begin{align*}
   2 \text{ mole Fe} \quad &55.85 \text{ g Fe} = 111.7 \text{ g Fe} \\
   3 \text{ mole O} \quad &16.00 \text{ g O} = 48.00 \text{ g O}
   \end{align*}
   \]

4. **Add the mass of each element to determine the Molar Mass of the compound:**
111.7 g Fe + 48.00 g O = 159.7 g Fe$_2$O$_3$

5. Divide the mass of each element in the compound by the Molar Mass of the compound, and multiply by 100.

\[
\frac{111.7 \text{ g Fe}}{159.7 \text{ g Fe}_2\text{O}_3} \times 100 = 69.94\% \text{ Fe}
\]
\[
\frac{48.00 \text{ g O}}{159.7 \text{ g Fe}_2\text{O}_3} \times 100 = 30.06\% \text{ O}
\]

6. Check to make sure the %s add up to 100.

Do sample problem J on page 243.

Calculation of Simplest Formula

Simplest Formula or Empirical Formula: *Lowest whole number ratio of atoms or ions in a compound.*

For an ionic compound, the simplest formula is the only formula used. *For a molecular compound, we must also calculate the molecular formula, which may or may not be the simplest formula.*

**Example:** Calculate the Simplest formula of a compound which has the following % composition: 35.35 \% Cr, 38.07\% O, 26.58\% K

To Calculate Simplest formula from % composition: *Remember units of % composition are g element/ 100 g compound.*

1. Assume you have 100 g of the compound.

2. Use the % composition to calculate the # of grams of each element in the 100 grams of compound: In 100 g of the above compound, there are:

\[35.35 \text{ g Cr, } 38.07 \text{ g O, } 26.58 \text{ g K.}\]
3. Convert the masses of each element to moles using Molar Mass:

0.6798 moles Cr
2.379 moles O
0.6798 moles K

4. Divide each # of moles by the smallest number, to obtain a mole ratio. This mole ratio is the ratio of atoms in the compound, if it is a whole number ratio, it is the simplest formula.

\[ \text{K Cr O}_{3.5} \]

5. If you do not obtain a whole number ratio after step 4, multiply the ratio by small whole numbers, starting with 2,3,4,etc. until you obtain a whole number ratio.

\[ (\text{K Cr O}_{3.5}) \times 2 = \text{K}_2\text{Cr}_2\text{O}_7 \]

Do sample problem M on page 247

**Molecular Formulas**

Used for molecular compounds only.

Simplest formula is just a RATIO of the atoms in a molecule. A molecule usually has more atoms than is indicated by the smallest whole number ratio.

Molecular Formula – the exact number of each type of atom in a molecule. Molecular formula can be determined from the simplest formula if we know the molecular mass of the substance.

\[ (\text{Molar Mass of the Simplest Formula}) \times \text{X} = \text{Molecular Mass} \]

Molecular Formula = Simplest Formula \times \text{X}

Sample Problem N on p. 248:
Practice Problem 2. on page 249:

**Formula of a Hydrate**

**Hydrate:**

**Anhydrous Compound:**

**Waters of Hydration:**

**Example Problem**

Washing soda is a hydrate of sodium carbonate. A student wanted to determine how many waters of hydration are in a formula unit of the hydrate. The student took a sample of the hydrate and heated it to evaporate the water. To make sure all of the water had evaporated she massed it after heating it several times, until the mass was constant. The student obtained the following data:

Mass of the hydrate before heating: 2.714 g
Mass of the compound after heating to a constant mass: 1.006 g

Determine:
A. The mass % of water in the hydrate.

B. The formula of the hydrate.