Molecular Mass, Formula Mass & Moles

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The Physical Science Series
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Atomic Mass
The atomic mass is the mass of an atom of an element. It is measured in Atomic Mass Units.

- A proton (+) has a mass of 1 AMU
- A neutron (0) has a mass of 1 AMU
- An electron (0) has a mass of 0 AMU

The Atomic Mass of an element is shown on the period table. The mass is not a whole number because of isotopes.

Note: Drawings not to scale. Protons and neutrons are 1836 x more massive than electrons.
Molecular Mass (Covalent)

Covalent compounds share valence electrons to form molecules.

Molecular Mass is total mass of all the atoms in a molecule.

The molecular mass of water is 18.015 AMU.
2 Hydrogen (2 x 1.008 AMU) + 1 Oxygen (15.999 AMU)
To determine the molecular mass of a covalent compound, simply add the masses of all the atoms in the compound.

Boron Trifluoride
\[
\text{BF}_3: \quad B \times 1 = 10.82 \text{ AMU} \\
\quad + \quad F \times 3 = 57.000 \text{ AMU} \\
\quad = \quad 67.82 \text{ AMU}
\]

Phosphorus Pentachloride
\[
\text{PCl}_5: \quad P \times 1 = 30.974 \text{ AMU} \\
\quad + \quad Cl \times 5 = 177.265 \text{ AMU} \\
\quad = \quad 208.239 \text{ AMU}
\]

Sig Figs in + and - Round your answer to the least precise mass you are adding.
Formula Mass
Formula Mass (Ionic)

Formula mass is the mass of a unit cell in an ionic compound.

A unit cell is composed of ions in the ratio of the formula.

The formula mass of salt is 58.448 AMU.
1 Sodium cation (22.991 AMU) + 1 Chloride anion (35.457 AMU)

Calculate formula mass the same way you do molecular mass.
Moles
**Moles**

**Number**: A mole is an amount of a substance equal to a quantity of 602 sextillion particles (atoms, molecules, or unit cells). Written as $602,000,000,000,000,000,000,000 = 6.02 \times 10^{23}$

**Mass**: A mole is the atomic mass (element), molecular mass (covalent compound) or the formula mass (ionic compound) of a substance expressed in grams.

Since atoms, molecules and unit cells are extremely small, a mole is a **huge number** of particles but a **small mass**.

**Mass of a sample ≠ number of particles in the sample!**
# The Mole: Number vs. Mass

<table>
<thead>
<tr>
<th>One mole of:</th>
<th>Number of particles</th>
<th>Molar Mass</th>
</tr>
</thead>
<tbody>
<tr>
<td>C</td>
<td>$6.02 \times 10^{23}$ atoms</td>
<td>12.011 g</td>
</tr>
<tr>
<td>CF$_4$</td>
<td>$6.02 \times 10^{23}$ molecules</td>
<td>88.01 g</td>
</tr>
<tr>
<td>Al(OH)$_3$</td>
<td>$6.02 \times 10^{23}$ unit cells</td>
<td>78.00 g</td>
</tr>
<tr>
<td>N$_2$O$_5$</td>
<td>$6.02 \times 10^{23}$ molecules</td>
<td>108.011 g</td>
</tr>
<tr>
<td>Cu</td>
<td>$6.02 \times 10^{23}$ atoms</td>
<td>63.546 g</td>
</tr>
<tr>
<td>Fe$_2$O$_3$</td>
<td>$6.02 \times 10^{23}$ unit cells</td>
<td>159.70 g</td>
</tr>
<tr>
<td>SiO$_2$</td>
<td>$6.02 \times 10^{23}$ molecules</td>
<td>60.09 g</td>
</tr>
</tbody>
</table>
Mole Conversions
Suppose you are given 70.0 grams of salt and asked to calculate the number of unit cells in the sample. How would you do it?

1) Calculate the molar mass of sodium chloride.

\[
\text{Na} = 22.991 \text{ g/mol} \\
\text{Cl} = +35.457 \text{ g/mol} \\
\text{NaCl Molar Mass} = 58.448 \text{ g/mol}
\]

2) Use DA to convert the given mass to unit cells.

\[
\frac{70.0 \text{ g}}{58.448 \text{ g}} \cdot \frac{1 \text{ mole}}{1 \text{ mole}} \cdot \frac{6.02 \times 10^{23} \text{ unit cells}}{1 \text{ mole}} = 7.21 \times 10^{23} \text{ unit cells}
\]
Mole Conversions

Whenever you are doing conversions (using DA) always start with your given (what you are given to convert).

One mole always shows up in your conversion factor. If you are given mass or particles, you must first convert it to moles.

The only two conversion factors you will use contain either:
- 1 mole and molar mass
- 1 mole and 602 sextillion atoms, molecules or unit cells.

Given Mass (grams) → Moles (moles) → Number of Particles (atoms, molecules, unit cells)
Converting mass to # of particles

Suppose you are given 8.63 grams of sulfur and asked to calculate the number of sulfur atoms in the sample.

Here is how you would do it:

Given Mass (grams) → Moles (moles) → Number of Particles (atoms, molecules, unit cells)

Answer: $1.620 \times 10^{23}$ sulfur atoms
You are told that a sample of copper(II) nitrate contains $3.48 \times 10^{24}$ unit cells and asked to calculate the mass of the sample.

Here is how you would do the calculation:

**Formula:** $\text{Cu(NO}_3\text{)}_2$

**Molar mass:** 187.56 g/mol

Answer: 1084 grams of copper(II) nitrate
You are told you have 17.90 moles of water. What is the mass and number of molecules in the water?

Formula: $\text{H}_2\text{O}$  \hspace{1cm} Molar Mass = 18.015 g/mol

\[ \frac{17.90 \text{ moles of water}}{1} \times 18.015 \text{ g/mole} = 322.5 \text{ g} \]

\[ \frac{17.90 \text{ moles of water}}{1} \times 6.02 \times 10^{23} \text{ molecules/mole} = 1.078 \times 10^{25} \text{ molecules} \]
Molecular and Formula masses measure the masses of tiny individual molecules and unit cells - too small to see. The unit for these measurements is the **AMU** (Atomic Mass Unit).

Molar mass is the mass of 602 sextillion atoms, molecules or unit cells. The unit for molar mass and given mass is **grams**.

Elements are composed of **atoms**.

Covalent compounds are composed of **molecules**.

Ionic compounds are composed of **unit cells** (aka formula units).
Mole Calculations:

**Given mass** is the mass you are given - it can be any amount.

**Molar mass** is the mass of one mole of the substance. It is calculated by adding up the atoms/ions and expressing in grams.

Because you either begin with moles or convert directly to moles, **1 mole** shows up in every conversion factor.

Rules for Sig figs:
- Round to **least precise** when calculating your molecular/formula or molar mass.
- Round to **least number of sig figs** when using DA.
What is October 23?

National Mole Day!

Everything and Everywhere Chemistry is