Module 3: Representing Amounts of Substances

The Mole and Molar Mass

Fundamentals of Chemistry Open Course

Learning Objectives | Module 3

1. Explain the significance of Avogadro's number and why the value $6.022 \times 10^{23}\ \text{mol}^{-1}$ is a convenient definition of the mole.

2. Use average atomic masses to calculate the molar mass of a substance with given chemical formula.

3. Apply molar mass to determine amount of substance from mass and *vice versa*.

4. Define mass density and molar volume; apply them in calculations.

5. Visualize liquid solutions at the submicroscopic level; identify the components of a solution.

6. Define concentration and recognize common units of concentration.

7. Define molarity and apply it to calculate amount of solute from volume of a solution and *vice versa*.

8. Recognize quantities in the ideal gas law and their associated units.

9. Apply the ideal gas law to calculate the amount of a gas from pressure, volume, and temperature.

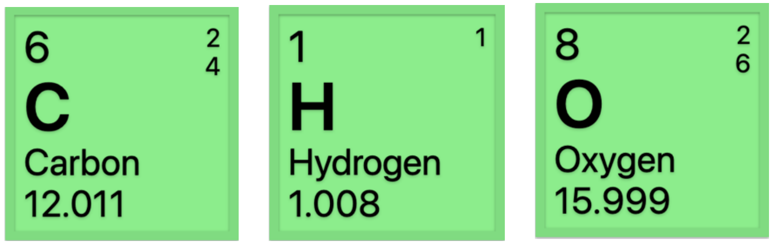
How Much Does a Molecule Weigh?

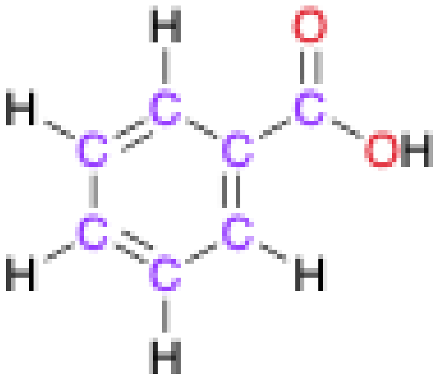
- The **atomic mass unit (u, amu)** is defined as one twelfth the mass of a carbon-12 atom.
  - One carbon-12 atom weighs 12.00 u; other isotopic masses are measured with respect to this standard.
  - One atomic mass unit is equivalent to 1.661×10^{-24} g.

- **Average atomic masses** on the periodic table are averages of all known isotopes of an element, weighted by their relative abundance.

- Macroscopic samples of chemical substances, which contain a *very* large number of atoms, always reflect average atomic masses.

- The **formula mass** of a substance is the sum of the average atomic masses of all atoms in the chemical formula.





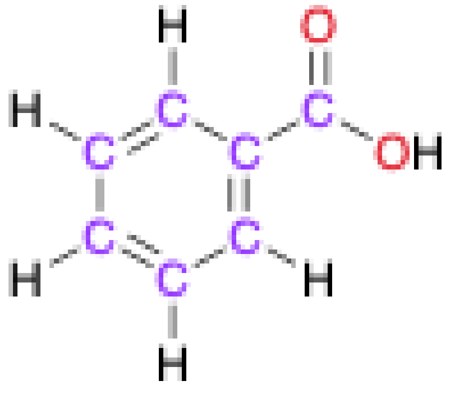
benzoic acid
$C_7H_6O_2$

$$7(12.011\ \mathrm{u}) + 6(1.008\ \mathrm{u}) + 2(15.999\ \mathrm{u}) = 123.123\ \mathrm{u}$$

Using Mass to Count Molecules

- If we know the mass of a compound and the mass of one of the particles in the sample, we can use proportional reasoning to determine the *number* of particles in the sample.

Example. How many molecules are in a sample of benzoic acid weighing 5.00 grams?



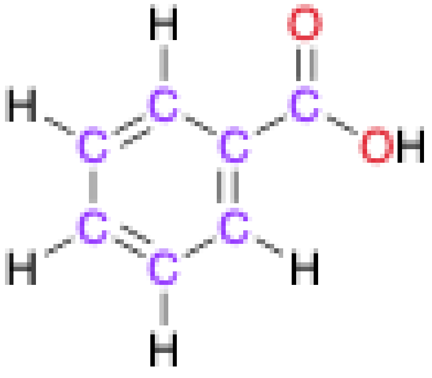
benzoic acid
$C_7H_6O_2$, 122.123 u

- Macroscopic samples of compounds contain *huge* numbers of molecules.
  We need a very large counting unit for molecules.

The Mole and Avogadro’s Number

- The **mole (mol)** is the number of particles in a substance with a mass in grams equal numerically to the mass of a single particle in atomic mass units. This is a highly convenient definition!

- For example, 1 mole of benzoic acid has a mass of 122.123 grams.



benzoic acid
$C_7H_6O_2$, 122.123 u

1 mole
122.123 g

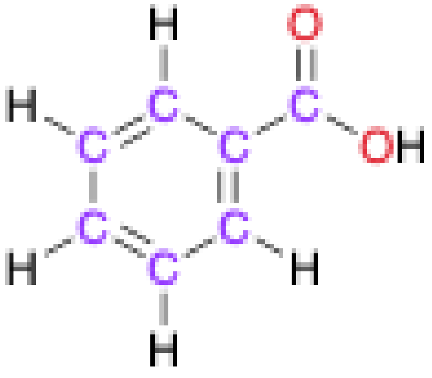
- The mole is just a counting unit—a representation of a count of atoms, molecules, ions, photons, reactions, etc. Number of particles in a substance (measured in moles) is referred to formally as **amount of substance**.

- Using the mole, we can easily keep track of numbers of particles using macroscopic masses in grams.

The Mole and Avogadro's Number | 1 of 2

- One mole corresponds to a fixed number of particles: **Avogadro's number N_A**.

- The number of particles in 1 mole is equal to the mass of 1 mole of a substance in grams divided by the formula mass *in grams*. For example, using benzoic acid,



benzoic acid
$C_7H_6O_2$, 122.123 u

1 mole
122.123 g

$$N_A = \frac{122.123\ \text{g}}{122.123\ \text{u}\left(\frac{1.661\times 10^{-24}\ \text{g}}{1\ \text{u}}\right)} = 6.022 \times 10^{23}$$

- Thanks to the definition of the mole, Avogadro's number is the same for all substances.

- Avogadro’s number has units of “[anything] per mole,” often represented as mol^{-1} or /mol.

$$N_A = 6.022 \times 10^{23}\ \mathrm{mol}^{-1}$$

- The objects to which a count refers are typically clear from the context: molecules for a molecular compound, formula units for an ionic compound, reaction occurrences for thermodynamic quantities, etc.

- However, when working with mole units, including a chemical formula after “mol” is strongly advised. This formula answers the vital question: “moles of *what?*”

- Avogadro’s number is used to “convert” from an absolute count to units of moles and *vice versa*.

Using Moles to Count Particles

- The mole is significantly more convenient than an absolute count when working with macroscopic substances or reactions.

Example. We previously calculated the absolute count of molecules in a 5.00-gram sample of benzoic acid. What is this count in units of moles?

Example. What is the absolute count of formula units in 2.50 mol $CaCO_3$?

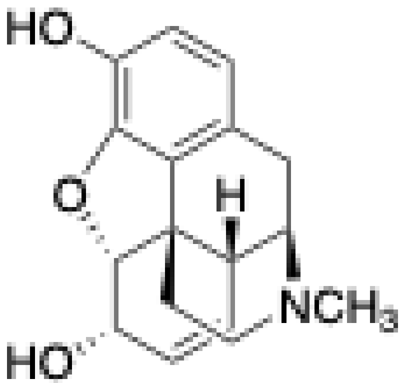
Molar Mass: How Much Does a Mole Weigh?

- From the definition of the mole, it follows that the **molar mass** of any substance in grams (for 1 mole) is numerically equal to the formula mass of the substance in atomic mass units (for 1 particle).

- Thus, we can apply our known method for calculating formula mass to determine molar mass as well.
  - Average atomic masses on the periodic table are also molar masses (just a change of units).
  - The molar mass of a substance is the sum of the molar masses of all atoms in its chemical formula.

- Molar mass has units of grams per mole (g/mol or g mol^{-1}).

Example. What is the molar mass of morphine, which has the molecular formula $C_{17}H_{19}NO_3$?



morphine
$C_{17}H_{19}NO_3$

Lingering Questions

- Can the mole and Avogadro’s number be defined with respect to a mass unit other than the gram?
- What about substances dissolved in solutions and other substances that we can’t weigh?
  How do we count particles of these substances?
- How are moles used in planning and carrying out chemical reactions?