Topic 1.1

Moles and Molar Mass

Learning Objectives

- Describe the mole concept and Avogadro's number.
- Define the atomic mass unit and how it is used to calculate the mass of an element or molecule.

Topic Questions

- What units are used to measure the mass of atoms and ions?
- What unit conversion is needed to measure the mass of atoms and ions on a large scale?
- What is the relationship between a mole and Avogadro's number?

1.1.01 The Mole Concept and Avogadro's Number

[SPQ-1.A.1 SPQ-1.A.2]

In a sample of a substance, the individual particles (ie, atoms, ions, or molecules) are too small to be seen and counted. Therefore, to find the number of particles in a sample, a relationship between mass and number of particles must be used. This relationship is provided by **Avogadro's number** (6.022×10^{23}), which is the number of atomic mass units (amu) equal to a mass of exactly 1 gram.

$$6.022 \times 10^{23} \text{ amu} = 1.000 \text{ gram}$$

This is a very large number, so a counting set called a **mole** is used. A mole of a substance contains Avogadro's number of particles, or 6.022×10^{23} particles, of that substance (Figure 1.1).

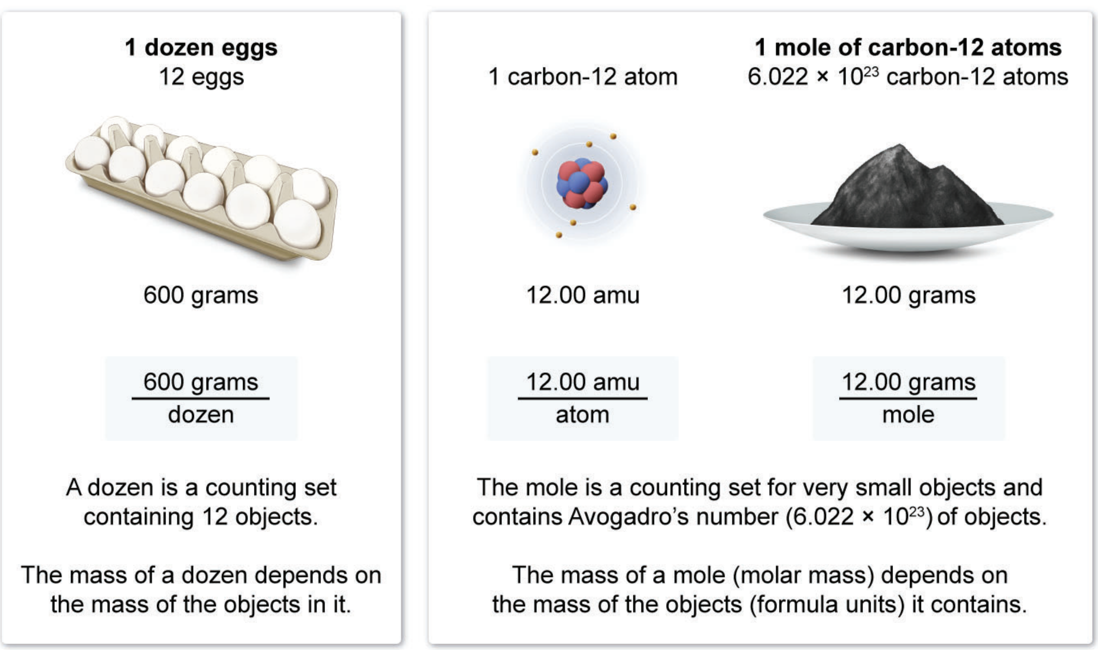

Figure 1.1 The mole concept.

The mass present in 1 mole of an element is known as the element's **molar mass**. On the periodic table, the mass listed for each element represents both the molar mass and **atomic mass** of the element. The molar mass is measured in grams per mole (g/mol), while atomic mass is measured in atomic mass units (amu). Therefore, an element's atomic mass and molar mass are numerically the same but have different units.

A mole can be used to find the mass and number of particles of a substance, as shown in Figure 1.2.

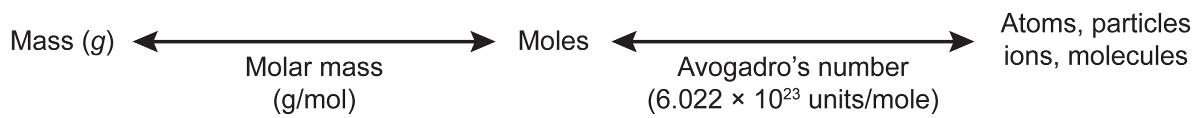

Figure 1.2 Conversion pathway between mass, moles, and particles.

For example, the molar mass of copper (Cu) is 63.55 g/mol, so 1 mole of Cu is 63.55 grams, and one mole of Cu contains 6.022×10^{23} atoms of Cu.

$$1\ \cancel{\text{mol}}\ \text{Cu} \times \frac{63.55\ \text{g Cu}}{1\ \cancel{\text{mol}}} = 63.55\ \text{g Cu}$$

$$1\ \cancel{\text{mol}}\ \text{Cu} \times \frac{6.022 \times 10^{23}\ \text{atoms Cu}}{1\ \cancel{\text{mol}}} = 6.022 \times 10^{23}\ \text{atoms Cu}$$

The reverse is also true: dividing the mass of a pure sample by its molar mass or dividing the number of particles by Avogadro's number gives the number of moles of the substance in the sample.

$$63.55\ \cancel{\text{g}}\ \text{Cu} \times \frac{1\ \text{mol Cu}}{63.55\ \cancel{\text{g}}\ \text{Cu}} = 1\ \text{mol Cu}$$

$$6.022 \times 10^{23}\ \cancel{\text{atoms}}\ \text{Cu} \times \frac{1\ \text{mol Cu}}{6.022 \times 10^{23}\ \cancel{\text{atoms}}\ \text{Cu}} = 1\ \text{mol Cu}$$

1.1.02 Atomic Mass Units

[SPQ-1.A.3]

The number of **protons** in an atom determines what type of atom it is (ie, which **element**). For example, every atom with 1 proton is hydrogen (H), whereas every atom with 2 protons is helium (He). **Isotopes** are atoms of the same element (ie, atoms with the same number of protons) that have different atomic masses due to differences in the number of **neutrons**, as illustrated in Figure 1.3.

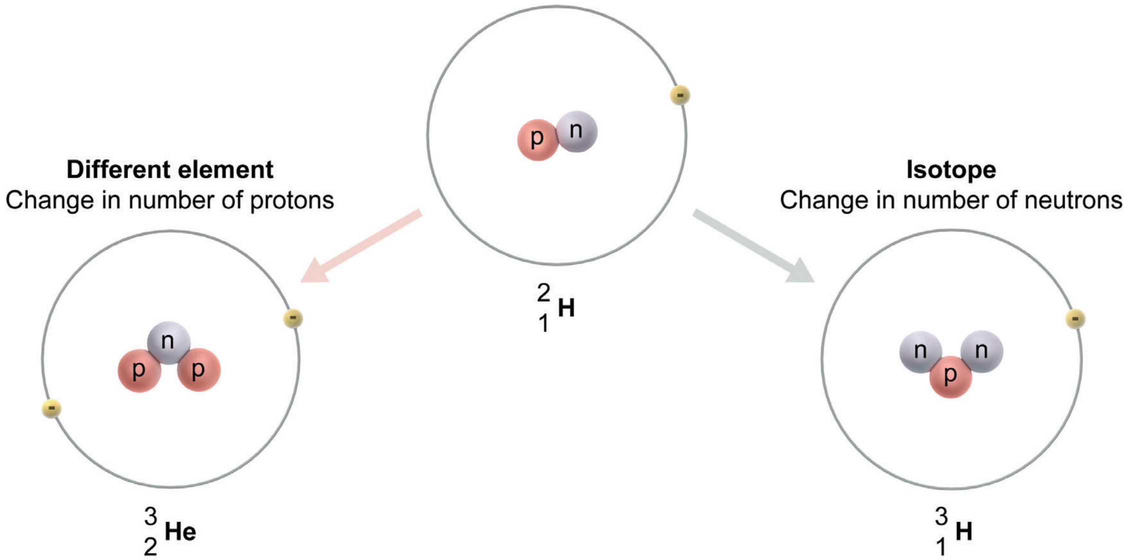

Figure 1.3 The relationship between isotopes and different elements.

Because all atoms are made of protons and neutrons, the isotope carbon-12 is a convenient natural standard to use for measuring the atomic masses of the other elements on the periodic table. The **atomic mass unit** (u or amu) is a unit of mass defined as one-twelfth the mass of a neutral, unbound atom of carbon-12 (Figure 1.4).

$$\frac{\text{Mass of } ^{12}_{6}\text{C atom}}{12} = 1 \text{ amu}$$

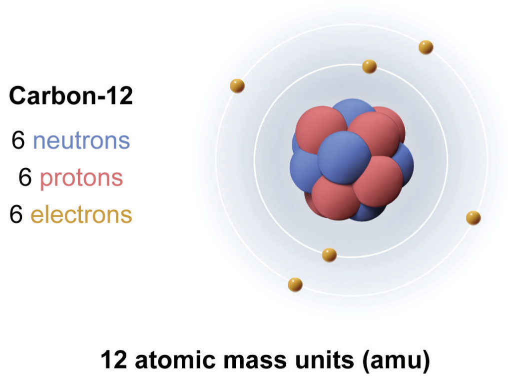

Figure 1.4 The atomic mass unit (amu) is one-twelfth the mass of a carbon-12 atom.

The atomic mass unit provides a useful unit to measure the masses of different atoms. However, because the amu is very small, the masses of large quantities of atoms are measured on the **gram scale** (ie, macroscale). As discussed in Sub-Topic 1.1.01, to relate small-scale masses (such as amu) to large-scale masses (such as grams), the mole is used. Remember that one mole represents 6.022×10^{23} items (ie, Avogadro's number):

$$1 \text{ mole} = 6.022 \times 10^{23} \text{ items}$$

This number is also the number of amu equal to 1.00 g of mass.

$$6.022 \times 10^{23} \text{ amu} = 1.00 \text{ g}$$

Therefore, because each carbon-12 atom has a mass of 12 amu, the total mass of a mole of carbon-12 atoms is exactly 12.00 grams:

$$1.00 \cancel{\text{mol}}\ {}^{12}\text{C atoms} \times \left(\frac{6.022 \times 10^{23}\ {}^{12}\text{C } \cancel{\text{atoms}}}{1 \cancel{\text{mol}}\ {}^{12}\text{C atoms}}\right) \times \left(\frac{12 \cancel{\text{amu}}}{1\ {}^{12}\text{C } \cancel{\text{atom}}}\right) \times \left(\frac{1.00 \text{ g}}{6.022 \times 10^{23} \cancel{\text{amu}}}\right) = 12.00 \text{ g}$$

Based on this calculation, the **atomic mass** and the **molar mass** of an element have the same numerical value but different units. In the same way, the **molecular mass** and the molar mass of a **compound** also have the same numerical value but different units.

The mass of an individual molecule (in amu) is found by adding together the atomic masses (from the periodic table) for each atom in the molecule. For example, the mass of a *single* CH_2Cl_2 molecule is:

$$(\text{C mass}) + 2(\text{H mass}) + 2(\text{Cl mass}) = \text{CH}_2\text{Cl}_2 \text{ mass}$$

$$12.01 \text{ amu} + 2(1.01 \text{ amu}) + 2(35.45 \text{ amu}) = 84.93 \text{ amu}$$

Therefore, the mass of 1 mole of CH_2Cl_2 on the gram scale is 84.93 g (Figure 1.5).

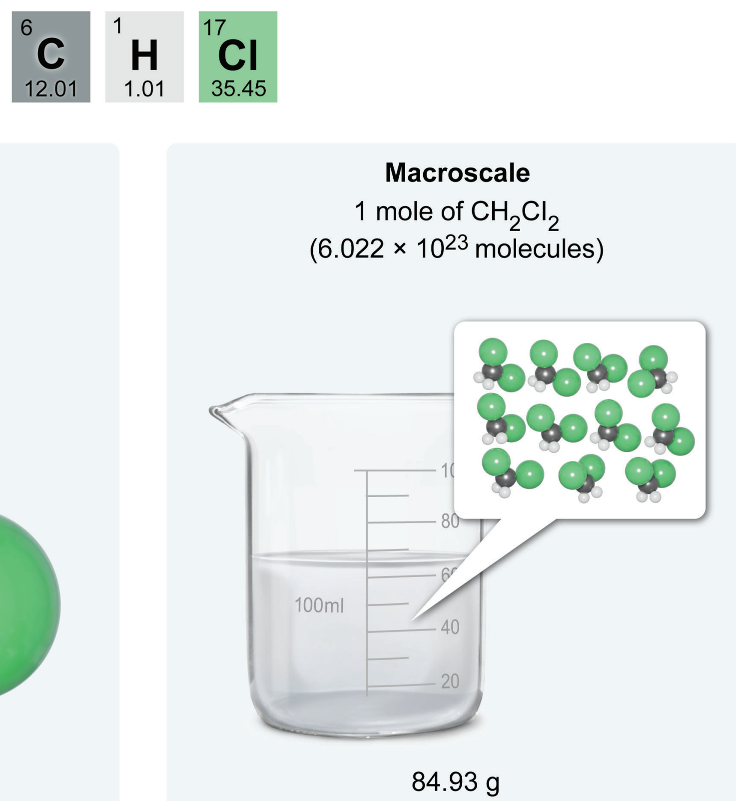

Figure 1.5 Atomic scale versus macroscale (ie, gram-scale) masses of dichloromethane (CH_2Cl_2).

Topic 1.1 Moles and Molar Mass

Check for Understanding Quiz

1. How many atoms are in 0.050 moles?

   A. 8.3×10^{-26} atoms

   B. 5.0×10^{-2} atoms

   C. 3.0×10^{22} atoms

   D. 6.0×10^{23} atoms

2. Which of the following terms refers to the amount of mass that is present in 1 mole of a substance?

   A. Atomic mass

   B. Avogadro's number

   C. Molar mass

   D. Molecular weight

3. What is the molar mass of glucose, $C_6H_{12}O_6$?

   A. 4.75 g/mol

   B. 29.02 g/mol

   C. 180.16 g/mol

   D. 240.08 g/mol

Note: Answers to this quiz are in the back of the book (appendix).