Activity 03-3

Mole and Molar Mass

Why?

To keep track of the huge numbers of atoms and molecules in samples that are large enough to see, chemists have established a unit of counting called the *mole* (abbreviated mol) and a unit of measure called the *molar mass*, which has units of g/mol. By using the idea of a *mole* and *molar mass*, you will be able to count out specific numbers of atoms or molecules simply by weighing them. This capability is necessary in understanding chemical reaction equations, conducting research in chemistry and biology, and applying chemistry in technology and the health sciences.

Learning Objectives

- Understand the relationship between the mole and Avogadro's number
- Understand the meaning of the molar mass of a substance
- Recognize that the molar mass is an average of all the isotopic masses of an element

Success Criteria

- Quickly convert between the number of atoms, moles, and the mass of a sample by using Avogadro's number and the molar mass appropriately
- Calculate the molar mass from isotopic abundances and isotopic masses

Prerequisites

- **Activity 01-2:** *Unit Analysis*
- **Activity 01-3:** *Significant Figures in Measurements and Calculations*
- **Activity 02-1:** *Atoms, Isotopes, and Ions*

Model 1: A Mole is a Counting Unit

1 pair of objects = 2 objects

1 dozen objects = 12 objects

1 gross of objects = 144 objects

1 mole of objects = 6.02214×10^{23} objects

Key Questions

1. How many pencils are there in a dozen pencils? *12*

2. How many pencils are there in a gross of pencils? *144*

3. How many pencils are there in a mole of pencils? *6.02214×10^{23}*

4. How many atoms are there in a dozen atoms? *12*

5. How many atoms are there in a gross of atoms? *144*

6. How many atoms are there in a mole of atoms? *6.02214×10^{23}*

7. In what way are the meanings of the terms *pair*, *dozen*, *gross*, and *mole* similar?

   They all refer to a number of objects and are used to group objects to make counting them easier.

   In what way are the meanings different?

   The numbers they represent differ. A mole is a much larger number because it is typically used to count very small objects.

Information

The number of objects in a mole (6.02214×10^{23}) is so important in chemistry that is given a name. It is called *Avogadro's number*, which has units of objects /mol.

Avogadro's number is determined by the number of carbon atoms in exactly 12 g of pure carbon-12.

The *molar mass* is the mass of a mole of objects. It has units of g/mol.

1 amu = 1.66054×10^{-24} g

Exercises

1. A single carbon-12 atom has a mass of 12 amu by definition of the atomic mass unit. Convert 12 amu to grams, and then calculate the mass in grams of a mole of carbon-12 atoms.

   $$12\ amu\,(1.66054 \times 10^{-24}\ g / amu) = 1.99265 \times 10^{-23}\ g / atom$$
   $$(1.99265 \times 10^{-23}\ g / atom) \times (6.02214 \times 10^{23}\ / mol) = 12.0000\ g / mol$$

2. A single oxygen-16 atom has a mass of 15.9949 amu. Convert this mass to grams, and then calculate the mass in grams of a mole of oxygen-16 atoms.

   $$15.9949\ amu\,(1.66054 \times 10^{-24}\ g / amu) = 2.65602 \times 10^{-23}\ g / atom$$
   $$(2.65602 \times 10^{-23}\ g / atom) \times (6.02214 \times 10^{23}\ / mol) = 15.9949\ g / mol$$

3. Based on your results for Exercises 1 and 2, identify the relationship between the numerical values of the mass of an atom in amu and the molar mass in g/mol.

   They are numerically the same because the conversion factor from amu to g is just the inverse of Avogadro's number!

Model 2: Beads in a Jar—Average Mass of a Mixture of Objects

Natural samples of most elements are mixtures of different isotopes. The mass of Avogadro's number of atoms in such a sample is not the molar mass of a single isotope but is rather an abundance weighted average of the masses of all the isotopes for that element. Exploration of **Model 2** will guide you in determining these average molar masses.

Table 1

Bead Color	Mass of a Single Bead	Number in the Jar	Percent Abundance
red	2.0 g	50	50%
blue	2.5 g	30	30%
yellow	3.0 g	20	20%

Key Questions

8. Which color bead has the largest mass? *yellow*
9. Which color bead is present in the largest number? *red*
10. Is the average mass of a bead in the jar equal to 2.5 g, which is [(2.0 + 2.5 + 3.0) ÷ 3]? Explain.

    No, the average mass is not 2.5 g because there are unequal numbers of red, blue, and yellow beads.

11. Is the average mass of a bead in the jar greater than or less than 2.5 g? Explain, without doing a calculation; just examine the information in Table 2.1.

    The average mass is less than 2.5 g, because there are many more of the lighter, red (2.0 g) beads than the heavier, yellow (3.0 g) beads.

12. How can you calculate the average mass of a bead in the jar using the percent abundance given in Table 2.1? Provide an explanation, then do the calculation.

    Multiply the mass of each bead by the fraction present and add the results for all three colors:

    $$(0.5 \times 2.0\ g) + (0.3 \times 2.5\ g) + (0.2 \times 3.0\ g) = 2.35\ g$$

13. How can you calculate the molar mass of the beads using the average mass that you calculated in the previous question? Provide an explanation, then do the calculation.

    Multiply the average mass by Avogadro's number:

    $2.35\ g \times (6.02214 \times 10^{23}\ /mol)$

    $= 1.42 \times 10^{24}\ g/mol$

EXERCISES

4. Using your calculation of the molar mass of beads as a guide, show how to determine the molar mass of boron from the data given in Table 2.2 below. Remember, molar mass has units of g/mol and 1 amu = 1.66054×10^{-24} g.

Table 2

Isotope	Atomic Mass (amu)	Percent Abundance
boron-10	**10.0129**	**19.78%**
boron-11	**11.0093**	**80.22%**

The average atomic mass requires calculating weighted average of the masses.

Average atomic mass = 0.1978 (10.0129 amu) + 0.8022 (11.0093 amu)

Average atomic mass = 10.812 amu

The molar mass (in g/mol) will be the same number as the average atomic mass (in u), as noted in Exercise 3, so the molar mass of boron is 10.812 g/mol.

5. Compare the number you calculated for the molar mass of boron in Exercise 4 with the number given below the symbol for boron in the Periodic Table. From this comparison, identify the information that is provided by the numbers just below the atomic symbols in the Periodic Table.

The numbers are the same. All the numbers on the Periodic Table are average atomic masses in amu or the molar masses in g/mol. The other number given in the Periodic Table is the atomic number, which is the number of protons in the atom.

6. Calculate the number of atoms in exactly 2 moles of helium.

$2 \text{ mol} \times (6.02214 \times 10^{23}/\text{mol}) = 12.04428 \times 10^{23}$

7. Calculate the number of moles corresponding to 2.007×10^{23} atoms of helium.

$$\left(\frac{2.007 \times 10^{23}}{6.02214 \times 10^{23}}\right) = 0.3333 \text{ mol}$$

8. Calculate the mass in grams of 2.5 moles of argon. The molar mass of argon is 39.95 g/mol.

2.5 mol Ar × 39.95 g/mol Ar = 99.875 g Ar = 100 g Ar

Note 1: The units in the answer can be derived by carrying out the arithmetic operations on the units.

Note 2: The number of significant figures reported in the answer is derived from the number of significant figures in the numbers being multiplied. The two figure precision for moles of Ar limits the answer to two significant figures (1.0×10^2)

9. Calculate the number of moles in 75 g of iron. The molar mass of iron is 55.85 g/mol.

$$75 \text{ g Fe} \times \left(\frac{1 \text{ mol Fe}}{55.85 \text{ g Fe}}\right) = 1.3 \text{ mol Fe}$$

10. Calculate the number of atoms in 0.25 moles of uranium.

$0.25\ mol \times (6.02214 \times 10^{23}\ /mol) = 1.5 \times 10^{23}$

11. Calculate the mass of 12.04×10^{23} atoms of uranium. The molar mass of uranium is 238.0 g/mol.

$$\left(\frac{12.04 \times 10^{23}}{6.02214 \times 10^{23}}\right) = 238.0\ g/mol = 475.8\ g$$

Got It!

1. Identify the statement below that is correct and explain why it is correct:
   a) The molar mass of an element divided by Avogadro's number gives the average mass of an atom of that element in grams.
   b) The molar mass of an element divided by Avogadro's number gives the mass of one atom of that element in grams.

   a is correct because the molar mass is the mass of a sample consisting of all the natural isotopes of an element so dividing the molar mass by number of atoms in a mole gives the average mass of an atom.

   b is not correct because atoms of different isotopes in a natural sample have different masses.

2. If you have 1 g samples of several different elements, will the sample with the largest or smallest molar mass contain the fewest atoms? Explain.

   If the atoms are heavy, i.e. have a large molar mass, then few of them will be needed to add up to 1 g. So the sample with the largest molar mass will contain the fewest atoms.

   Mathematically 1g / molar mass = a small number of moles if the molar mass is large.

   Note: If students have trouble thinking about this situation, ask them which will have fewer particles, a pound of sand or a pound of coarsely crushed rock.

3. Write the units that result from the following mathematical operations.
   a) number of objects / Avogadro's number = *mol*
   b) moles × Avogadro's number = *gives the number of objects, no units*
   c) mass / molar mass = *mol*
   d) moles × molar mass = *g*

Problems

1. The atomic mass of ^{35}Cl is 34.971 amu and the atomic mass of ^{37}Cl is 36.970 amu. In a natural sample, 75.77% of the atoms are ^{35}Cl, and 24.23% are ^{37}Cl. Describe how you can calculate the molar mass of chlorine from these data, then calculate a value for the molar mass of chlorine.

   Calculate the weighted average.

   Average atomic mass = 0.7577 (34.971 amu) + 0.2423 (36.970 amu)

   Average atomic mass = 35.46 amu

   Molar mass = 35.46 g/mol

2. A mass of 32.0 g of oxygen reacts completely with 6.02214×10^{23} atoms of carbon.

   a) What is the ratio of moles of C to moles of O in the product?

   $$\frac{32.0 \text{ g O}}{(16.0 \text{ g/mol O})} = 2.00 \text{ mol O}$$

   $$\frac{6.02214 \times 10^{23} \text{ C atoms}}{6.02214 \times 10^{23}\text{/mol}} = 1.00000 \text{ mol C}$$

   C:O = 1:2

   b) Given that mass is conserved in a chemical reaction, what is the mass of the product produced?

   $$1.00000 \text{ mol C} \times \frac{12.01 \text{ g}}{\text{mol C}} = 12.01 \text{ g}$$

   32.0 g O + 12.01 g C = 44.0 g

   c) Is the product carbon monoxide, CO, or carbon dioxide, CO_2?

   CO_2 because the C:O ratio is 1:2.