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THE MOLE CONCEPT

Providing a name for a quantity of things taken as a whole is common in everyday life. Some examples are a dozen, a gross, and a ream. Each of these represents a specific number of items and is not dependent on the commodity. A dozen eggs, oranges, or bananas will always represent 12 items. In chemistry we have a unit that decribes a quantity of particles. It is called the **mole** (sometimes abbreviated as **mol**). A mole is 6.02×10^{23} particles. Technically, that's the number of carbon atoms found in exactly 12 grams of carbon-12. Since the atomic masses of all the elements' atoms are related to the mass of carbon-12, a mole is also the number of atoms found in the atomic mass of *any* element if it is *expressed in grams.* Keep in mind that the masses found on the Periodic Table for any element are actually weighted averages of all the isotopes that exist for that element (based on their relative natural abundances). Those masses, if expressed in atomic mass units (amu), represent just one average atom for that element. If, however, the value for mass was expressed in grams, that sample of the element

TIP

Just as a dozen is 12 units of an item, a mole is 6.02×10^{23} units of an item. It is called *Avogadro's number.*

would contain 6.02×10^{23} atoms of that element. This value is also known as **Avogadro's number** in honor of the Italian scientist whose hypothesis concerning the volumes of gases led to its determination. More on Avogadro's hypothesis will be discussed in an upcoming section on gas volumes and molar mass. You should recognize that Avogadro's number is very large because the items being counted (atoms) are very small. So 6.02×10^{23} atoms of most elements represent samples of atoms that are conveniently sized for working in the laboratory.

MOLAR MASS AND MOLES

The mass of a mole of particles is referred to as its **molar mass**. For moles of atoms, the atomic mass found on the Periodic Table for that element expressed in grams is the molar mass for that element. Some elements naturally exist as molecules, however. The molar mass of those elements takes into account the number of atoms in the molecule *in an additive manner*. Most elements are considered in a monatomic way (one atom). However, a few (hydrogen, nitrogen, oxygen, fluorine, chlorine, bromine, and iodine) are typically considered in a diatomic manner (two atoms) based on the way they are generally found to exist. This is not to say that you could not count moles of hydrogen atoms (H) as opposed to hydrogen molecules (H_2). You should, though, always be cognizant of the type of particle involved in any mole calculation.

➡️ **Example 1** _____

Determine the molar mass of silicon, nitrogen, and iron using the Periodic Table.

The atomic mass of silicon is 28.1 amu as found on the Periodic Table. Therefore, the molar mass of silicon is 28.1 g and represents 6.02×10^{23} *atoms* of silicon or 1 mole of silicon atoms.

The atomic mass of nitrogen is 14.0 amu as found on the Periodic Table. Nitrogen (N_2) is a diatomic element, however. It has a molar mass of 28.0 g, which represents 6.02×10^{23} *molecules* of nitrogen or 1 mole of nitrogen molecules. A mole of nitrogen (N) atoms would have a mass of 14.0 g if they were the appropriate particle to be considered in a given circumstance.

The atomic mass of iron is 55.8 amu. So iron has a molar mass of 55.8 g. This represents a sample of 6.02×10^{23} atoms of iron or 1 mole of iron atoms. Iron is typically not found in nature as a diatomic molecule.

Elements are just one type of substance for which the molar mass can be found. The molar masses of compounds can be found in a way similar to that of diatomic elements. Just add up the molar masses of the individual elements found in the compound based on the compound's formula.

➡️ **Example 2** _____

Find the molar mass of NH_3 (this is a molecular compound known as ammonia).

The molar mass of nitrogen is 14.0 g, and the molar mass of hydrogen is 1.0 g. Since there are 3 hydrogen atoms per molecule of ammonia, the molar mass of ammonia is 17.0 g and represents 6.02×10^{23} molecules of ammonia or 1 mole of ammonia molecules.

➡ Example 3 _____

Find the molar mass of $CaCO_3$ (this is an ionic compound known as calcium carbonate).

The molar masses of calcium, carbon, and oxygen are 40.1 g, 12.0 g, and 16.0 g, respectively. Since there are 3 oxygen particles per formula unit of calcium carbonate in addition to the single particles of calcium and carbon, the molar mass of calcium carbonate is 40.1 g + 12.0 g + (3)16.0 g or 100.1 g. This represents 6.02×10^{23} formula units (the particle for an ionic compound) or 1 mole of calcium carbonate.

➡ Example 4 _____

Find the molar mass of $CuSO_4 \cdot 5\,H_2O$ (this is a hydrated ionic compound known as copper(II) sulfate pentahydrate).

The molar masses of copper, sulfur, oxygen, and water are 63.5 g, 32.1 g, 16.0 g, and 18.0 g, respectively. Since there are 4 oxygen particles and 5 water molecules per formula unit in addition to the single particles of copper and sulfur, the molar mass of copper(II) sulfate pentahydrate is 63.5 g + 32.1 g + (4)16.0 g + (5)18.0 g or 249.6 g. This represents 6.02×10^{23} formula units or 1 mole of copper(II) sulfate pentahydrate.

MOLAR MASS AND GAS VOLUMES

In 1811, Amedeo Avogadro made a far-reaching scientific assumption that also bears his name. **Avogadro's Hypothesis** states that equal volumes of different gases contain equal numbers of particles at the same temperature and pressure. It means that under the same conditions, the number of molecules of hydrogen in a 1-liter container is exactly the same as the number of molecules of carbon dioxide (or of any other gas) in a 1-liter container even though the individual molecules of the different gases have different masses and sizes. Because of the substantiation of this hypothesis by much data since its inception, it is often referred to as **Avogadro's Law** and can be added to the list of gas laws discussed in Chapter 5. Avogadro's Law shows the relationship between the volume and the number of particles of a gas sample when the temperature and pressure are constant:

$$\frac{V}{n} = k$$

In other words, volume and the number of gas particles are directly related.

Because the volume of a gas may vary depending on the temperature and pressure, a standard is set for comparing gases. As stated in Chapter 5, the standard conditions of temperature and pressure (abbreviated STP) are 273 K and 1 atmosphere. Because the relationship between volume and number of particles of a gas is direct when the temperature and pressure are constant, the molar mass of a gas (which represents a *set* number of particles, namely 1 mole) occupies a *set* volume. The volume of 22.4 L is recognized as the molar volume of any gas at STP.

USING MOLAR MASS AND MOLAR VOLUME

Molar mass and molar volume are typically used as conversion factors to change quantities of reactants expressed as masses or as volumes to moles for use in stoichiometry problems via dimensional analysis.

➡ Example 1

Find the number of moles of silicon present in 4.30 g of silicon.

Recall that silicon is an element that is not diatomic. It has a molar mass of 28.1 g. So 28.1 g of silicon contains 1.00 mol of silicon atoms if significant figures are kept in mind.

Use dimensional analysis:

$$\frac{4.30 \text{ g Si}}{1} \times \frac{1.00 \text{ mol Si}}{28.1 \text{ g Si}} = 0.0153 \text{ mol Si}$$

Note that when using dimensional analysis, the given quantity is simply multiplied by a factor that is equal to the value 1 since the numerator and denominator in that factor are equal to each other. Using this method causes the magnitude of the given quantity not to change. What does change is the unit in which the quantity is expressed. The given units cancel out, leaving only the unit in the numerator of the factor to describe the quantity.

➡ Example 2

Find the number of moles of calcium carbonate in 0.750 g of calcium carbonate.

Recall that calcium carbonate is an ionic compound. It has a molar mass of 100.1 g. So 100.1 g of calcium carbonate contains 1.00 mol of calcium carbonate.

Use dimensional analysis:

$$\frac{0.750 \text{ g CaCO}_3}{1} \times \frac{1.00 \text{ mol CaCO}_3}{100.1 \text{ g CaCO}_3} = 0.00749 \text{ mol CaCO}_3$$

➡ Example 3

Find the number of moles of ammonia in 0.300 L of ammonia at STP.

Recall that ammonia is a gas at STP. Since the given quantity is supplied as a volume and not as mass, molar volume (not molar mass) should be used in the dimensional analysis equation:

$$\frac{0.300 \text{ L NH}_3}{1} \times \frac{1.00 \text{ mol NH}_3}{22.4 \text{ L NH}_3} = 0.0134 \text{ mol NH}_3$$

REMEMBER