The Mole Concept

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Masses of “molecules”

We can take the masses of individual atoms and add them together to obtain the mass of a molecule or formula unit.

Example: H_2O 18.02 amu

<u>Terms</u>

Average mass per atom – Atomic weight

Average mass per “molecule” – Molecular weight, Molecular mass, Formula weight, Formula mass

The Mole

The mole is a unit of measurement equal to

6.022×10^{23} “things” (to 4 sf)

just like there are 12 “things” in a dozen.

This number is called ***Avogadro’s number***.

Why 6.022×10^{23}? Let’s look at a couple of calculations.

If the molecular weight (MW) of water is 18.02 amu, how many molecules are in 18.02 g?

If the atomic weight (AW) of sodium is 22.99 amu, how many atoms are in 22.99 g?

Thus the mole is the number of "things" that makes the numerical value of the mass per atom or molecule in units of amu equal to the mass in grams.

The actual definition states it a little differently – the mole is the amount of substance that contains as many particles (atoms, molecules, formula units, ions, etc.) as there are atoms in exactly 12 g of the carbon-12 isotope. Our choice of the amu as 1/12 of the mass of a carbon-12 atom makes the number of atoms 6.022×10^{23}.

This gives us 2 new conversion factors!

1 mole = 6.022×10^{23} particles

1 mole = ____?_____ g (The Molar Mass)

The number of grams has the same numerical value as the mass per atom or the mass per "molecule".

Examples: 89 g of sodium = ____________ atoms

1.50 moles of Na = _____________g Na

7.689×10^{20} atoms of Na = ____________ moles of Na

Percentage Composition by Mass

1. Since a substance is composed of identical repeating units (formula units or molecules), the relative proportions of a unit are the same as the whole substance

What is the percentage composition of water?

2. This percentage has a meaning that can be applied many ways

- percentage per molecule
- % per group of molecules

How many grams of water does it take to contain 50.0 g of hydrogen?

How many grams of oxygen are in a mole of water?

Empirical formula

The simplest ratio of atoms in a molecule of a compound

This can be found from the percentage composition or any composition data.

Why?

Since a mole of any element contains the same number of atoms, a ratio of moles will be the same as a ratio of atoms.

To find the empirical formula

1. Find the number of moles of each element in the sample

2. Determine the simplest whole number ratio between them, this is the empirical formula.

   If the ratio at first does not give whole numbers, multiply each ratio factor by the smallest whole number that would convert all the numbers to integers.

Molecular Formula

The molecular formula is a multiple of the empirical formula

To obtain the molecular formula find how many empirical formula units are contained in 1 molecule

We need the following information

1. Empirical formula

2. Molecular Weight / Molar Mass

After finding the number of empirical formula units per molecule multiply each subscript in the empirical formula by this factor

Hydrates

Hydrates are solid ionic compounds with water molecules attached in a specific ratio; an ionic crystal with water trapped on the inside in a specific ratio

To obtain the formula find the number of moles of ionic compound and the number of moles of water

Then determine the ratio; the ratio of moles will be the same as a ratio of molecules

Equations and Chemical Reactions

A chemical equation summarizes a chemical process (chemical change or physical change)

Reactants Þ Products

Equations use chemical symbols and must be <u>balanced</u> to give quantitative info.

- Balancing based on Law of Conservation of Matter
- The phases of each substance are often indicated with *g, l, s, aq*

Balancing uses coefficients to make sure number of atoms of each element on reactant side = number of atoms of each element on the product side

Stoichiometry

Definition: The study of the quantitative relationships between reactants and products in a chemical reaction

The <u>balanced</u> equation gives a ratio of moles between reactants and products since a ratio of molecules is the same as a ratio of moles.

Example

How much water can be produced from the reaction of 4.0 g of hydrogen gas with an excess of oxygen gas?

1. Write a balanced equation
2. Identify what you are given and what you need to solve for
3. Convert (if needed) what you are given to moles
4. Use stoichiometric ratios from balanced equation as a conversion from moles of what you have to what you want
5. Convert (if needed) to desired units

Example

How much oxygen gas is needed to react completely 4.0 g of hydrogen gas? (water is the product)

Notice that the law of conservation of matter holds!

Limiting reactants

What if we are given two quantities of substance?

Work the problem twice! Choose the smaller answer as the actual amount required or produced. Whatever you start with to obtain this smaller answer is the limiting reactant.

Example

How much water is produced if we react 4.00 g of hydrogen gas with 16.0 g of oxygen gas?

Since we can only make 18.0 g of water from 16.0 g of oxygen, the amount of oxygen limits the production of water and is called the limiting reactant.

What is present at the end of the reaction?

Since we started with 4.00 g of hydrogen, 2.00 g are left. So. . . . 18.0 g of water and 2.00 g of hydrogen are present at the end of the reaction.

Percentage yield

The numbers above are based on a perfect reaction. Suppose in the problem above that some of the water is lost in the measurement. We measure the relative amount actually obtained using the percentage yield.

$$\% \text{ yield} = \frac{\text{Actual yield}}{\text{Theoretical yield}}(100)$$

What is the percentage yield in the previous example if 14.0 g of water are produced?