Chapter 3

Moles, Balancing Chemical
Equations, Stoichiometry

Question you MUST be able to Answer

How many grams of hydrogen gas can be prepared from 5.70 g of SrH_2 and 4.75 g of H_2O given the other product is strontium hydroxide?

What do we need to do?

1. Write a balanced chemical equation.

Writing and Balancing Chemical Equations

The most important reason for learning about moles is because it clarifies the amounts of substances taking part in a chemical reaction.

A ***chemical reaction*** is a statement using formulas that expresses the identities and quantities of the substances involved in a chemical or physical reaction.

Example: $H_2\,(g) + F_2\,(g) \rightarrow 2\,HF\,(g)$

Read 1 molecule H_2 reacts with 1 molecule F_2 to yield 2 molecules HF.

Note: If you start with 2 H atoms, you have to end up with 2 H atoms.

What is involved in writing a balanced equation?

Let's consider the chemical statement: Magnesium wire reacts with oxygen gas to yield powdery magnesium oxide.

1. Translate the statement.

$$\underbrace{Mg + O_2}_{\text{reactants}} \xrightarrow{\text{yields}} \underbrace{MgO}_{\text{product}}$$

2. Balance the atoms by writing ***balancing coefficients*** – that is, a numerical multiplier of all atoms in the formula that follow it.

$$\underline{2}\ Mg + \underline{\quad}\ O_2 \rightarrow \underline{2}\ MgO$$

3. Specify the states of matter: solid (s), liquid (l), gas (g) or aqueous solution (aq)

$$2\,Mg\,(s) + O_2\,(g) \rightarrow 2\,MgO\,(s)$$

Remember in balancing the equation, we cannot balance the O by changing MgO to MgO_2. And we cannot add other reactants or products to balance the equation. Also, always use the smallest whole-number coefficients.

Octane (C_8H_{18}) burns in the presence of oxygen to yield carbon dioxide and water vapor.

Sample Question

How many grams of hydrogen gas can be prepared from 5.70 g of SrH_2 and 4.75 g of H_2O given the other product is strontium hydroxide?

What do we need to do?

1. Write a balanced chemical equation.
2. Convert g to moles.

Why do we have to convert g to moles?

Reactions are between molecules, not grams.

Must convert mass to number of molecules.

4.75 g of water has 1.59 x 10^{23} molecules

What are moles and why do we use them?

The Mole

Chemists devised a unit called the ***mole*** to count chemical entities (i.e., atoms, ions, molecules or formula units) by weighing them.

Defining the Mole

The ***mole*** is the SI unit for amount of substance. It is defined as the amount of a substance that contains the same number of entities as there are atoms in exactly 12. g of carbon-12.

This number is called ***Avogadro's number***:

1 mole contains 6.022 x 10^{23} entities.

Example:

1 mole carbon-12 contains 6.022 x 10^{23} carbon-12 atoms

1 mole H_2O contains 6.022 x 10^{23} H_2O molecules

The atomic mass of an element expressed in amu is numerically the same as the mass of 1 mole of atoms of the element expressed in grams. A similar relationship holds for the mass of compounds.

1 Fe atom has a mass of

1 mole of Fe atoms has a mass of

1 molecule of H_2O has a mass of

1 mole of H_2O has a mass of

Determining Molar Mass

The ***molar mass*** ($\mathscr{M}$) of a substance is the mass per mole of its entities (i.e., atoms, ions, molecules or formula units).

Units of molar mass are grams / mole (g/mole).

1. ***Elements*** – molar mass of an element is simply the atomic mass found on the periodic table.

   Example: $\mathscr{M}$ of neon = 20.18 g/mol

   $\mathscr{M}$ of molecular oxygen, O_2 = 2 x 16.00 g/mol = 32.00 g/mol

2. ***Compounds*** – molar mass of a compound if the sum of the molar masses of the atoms of each element in the chemical formula.

   Example:

   $\mathscr{M}$ of SO_2 = 32.07 g/mol + $\overbrace{(2 \times 16.00 \text{ g/mol})}^{32.0000 \text{ g/mol}}$ =64.07 g/mol

Converting Between Amount, Mass and Number of Chemical Entities

1. ***Converting between amount and mass.***

If you know the moles of a substance, you can calculate the mass by multiplying by the molar mass:

$$\text{amount (moles)} \times \frac{\text{no. of grams}}{\text{1 mole}} = \text{mass of substance}$$

If you know the mass of a substance, you can calculate the number of moles by multiplying by 1 over the molar mass:

$$\text{mass (g)} \times \frac{\text{1 mole}}{\text{no. of grams}} = \text{amount (moles)}$$

2. ***Converting between amount and number.***

Similarly, we can use Avogadro's number to convert moles to the number of entities:

$$\text{Amount (moles)} \times \frac{6.022 \times 10^{23}\ \text{entities}}{\text{1 mole}} = \text{no. of entities}$$

Or we can do the reverse,

$$\text{No. of entities} \times \frac{\text{1 mole}}{6.022 \times 10^{23}\ \text{entities}} = \text{amount (moles)}$$

Now we can put it all together and convert mass to the number of entities:

$$\text{Mass(g)} \times \frac{\text{1 mole}}{\text{no. of grams}} \times \frac{6.022 \times 10^{23}\ \text{entities}}{\text{1 mole}} = \text{no. of entities}$$

How many grams of Ag are in 0.0342 mol Ag?

How many Fe atoms are in 95.8 g of Fe?

Sample Question

How many grams of hydrogen gas can be prepared from 5.70 g of SrH_2 and 4.75 g of H_2O given the other product is strontium hydroxide?

What do we need to do?

1. Write a balanced chemical equation.
2. Convert g to moles.
3. Convert moles of reactant to moles of hydrogen gas.

Stoichiometry means the study of the quantitative aspects of formulas and reactions.

Calculating Quantities of Reactants and Products

A balanced chemical equation is essential for all calculations involving amounts of reactants and products: if you know the number of moles of one substance, the balanced equation tells you the number of moles of all the others in the reaction.

<u>**Stoichiometrically Equivalent Molar Ratios from Balanced Equations**</u>

Let's consider the combustion of propane (C_3H_8). . .

C_3H_8 (g) + $5O_2$ (g) $\rightarrow$ $3CO_2$ (g) + $4H_2O$ (g) (balanced)

	<u>Equivalent Molar Ratio</u>
1 mol C_3H_8 reacts with 5 mol O_2	1 mol C_3H_8/5 mol O_2 **OR** 5 mol O_2/1 mol C_3H_8
1 mol C_3H_8 produces 3 mol CO_2	1 mol C_3H_8/3 mol CO_2 **OR** 3 mol CO_2/1 mol C_3H_8
1 mol C_3H_8 produces 4 mol H_2O	1 mol C_3H_8/4 mol H_2O **OR** 4 mol H_2O/1 mol C_3H_8

Other equivalent molar ratios:

5 mol O_2/3 mol CO_2 **OR** 5 mol O_2/4 mol H_2O **OR** 3 mol CO_2/4 mol H_2O

C_3H_8 (g) + $5O_2$ (g) $\rightarrow$ $3CO_2$ (g) + $4H_2O$ (g) (balanced)

1 mole of C_3H_8 reacts with 5 moles of O_2 to form 3 moles of CO_2 and 4 moles of H_2O

Let's determine the number of moles O_2 required if we know there are 10.0 moles H_2O produced. . .

$$\text{Moles } O_2 = 10.0 \text{ mol } H_2O \times \frac{5 \text{ mol } O_2}{4 \text{ mol } H_2O} = 12.5 \text{ mol } O_2$$

Approach for solving ANY stoichiometric problem

1. Write the balanced equation. *(You cannot solve this type of problem without a balanced equation.)*
2. Convert given mass (or number of entities) to moles.
3. Select appropriate equivalent molar ratio.
4. Convert moles given to moles desired substance.
5. Convert to mass or number of entities (if required).

Chlorine gas can be made in the laboratory by the reaction of hydrochloric acid and manganese (IV) oxide:

$4HCl\ (aq) + MnO_2\ (s) \rightarrow MnCl_2\ (aq) + 2H_2O\ (l) + Cl_2\ (g)$ [balanced]

When 1.82 mole of HCl reacts with excess MnO_2, (a) how many moles of Cl_2 form? (b) How many grams of Cl_2 form?

Sample Question

How many grams of hydrogen gas can be prepared from 5.70 g of SrH_2 and 4.75 g of H_2O given the other product is strontium hydroxide?

What do we need to do?

1. Write a balanced chemical equation.
2. Convert g to moles.
3. Convert moles of reactant to moles of hydrogen gas.
4. Determine which reactant produces the smaller amount of product.

Recipes

- 4 eggs
- 3 c flour
- 2 c sugar
- 1 tsp salt

If you have 1 doz eggs, 12 c of flour, 12 c of sugar, and 12 tsp of salt, how many complete cakes can you make?

Chemical Reactions that Involve Limiting Reactants

$2Cu_2S(s) + 3O_2(g) + 2C(s) \rightarrow 2SO_2(g) + 4Cu(s) + 2CO(g)$

100 g of Cu_2S will react O_2 from the atmosphere to form SO_2 until all Cu_2S is used up.

Cu_2S is the ***limiting reactant*** because the reaction stops once the Cu_2S is gone, no matter how much O_2 is present.

Solving Limiting Reactant Problems

In limiting reactant problems, the quantities of two (or more) reactants are given and we first determine which is the limiting reactant.

1. Write the balanced chemical equation.
2. Calculate how much product could be formed from the given quantity of each reactant.
3. The limiting reactant is the one that produced the ***least*** product.

Sample Question

How many grams of hydrogen gas can be prepared from 5.70 g of SrH_2 and 4.75 g of H_2O given the other product is strontium hydroxide?

Hydrazine (N_2H_4) and dinitrogen tetroxide (N_2O_4) ignite on contact to form nitrogen gas and water vapor. How many grams of nitrogen gas form when 1.00×10^2 g of N_2H_4 and 2.00×10^2 g N_2O_4 are mixed? How much of either reactant remains after the reaction?

What is the mass percent of carbon in glucose ($C_6H_{12}O_6$)? ($\mathcal{M}$ of $C_6H_{12}O_6$ = 180.16 amu) And what is the mass of carbon in 16.55 g of glucose?

The Importance of Mass Percent

For many purposes it is important to know how much of an element is present in a given amount of compound. In this section we find the composition of a compound in terms of mass percent…

Determining Mass Percent from a Chemical Formula

Each element in a compound constitutes its own particular portion of the compound's mass:

1. For a molecule (or formula unit) of a compound

Mass % of element **X** =

$$\frac{\overbrace{\text{Atoms of } \mathbf{X} \text{ in formula} \times \text{Atomic mass of } \mathbf{X} \text{ (amu)}}^{\text{Total mass of } \mathbf{X} \text{ in compound}}}{\text{Molecular (or formula) mass of compound (amu)}} \times 100\%$$

2. For a mole of compound

Mass % of element **X** =

$$\frac{\overbrace{\text{Moles of } \mathbf{X} \text{ in formula} \times \text{Molar mass of } \mathbf{X} \text{ (g/mol)}}^{\text{Total mass of } \mathbf{X} \text{ in 1 mole of compound}}}{\text{mass of 1 mol of compound(g)}} \times 100\%$$

Determining the Formula of an Unknown Compound

Empirical Formulas

Empirical formulas express the relative number of atoms of each element in a compound. An analytical chemist investigating an unknown compound will often go through a three step process to determine its empirical formula:

1. Decompose it to find the mass of each component element,
2. Mathematically convert these masses to moles and write a preliminary formula.
3. Mathematically convert the moles into whole-number subscripts to find the empirical formula.
   - Divide each subscript by the smallest subscript, and
   - (If necessary) multiply through by the smallest integer that turns all subscripts into integers.

Elemental analysis of a sample of an ionic compound showed 2.82 g Na, 4.35 g Cl and 7.83 g O. Calculate the empirical formula.

Molecular Formulas

If we know the molar mass of a compound, we can use the empirical formula to obtain the molecular formula.

1. We simply divide the molar mass by the *empirical formula mass* to obtain the whole-number multipler.

2. Multiply the empirical formula subscripts by the whole-number multiplier to obtain the molecular formula!

Sample Problem 3.10 – Elemental analysis shows that lactic acid ($\mathcal{M}$ 90.08 g/mol) contains 40.0 mass % C, 6.71 mass % H and 53.3 mass % O.

(a) Determine the empirical formula of lactic acid.

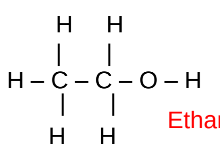
(b) Determine the molecular formula of lactic acid.

Isomers

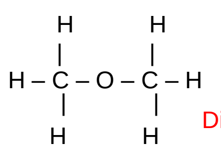
Recall, a molecular formula tells *the actual number and type of each atom in a molecule of a compound*. Yet different compounds can have the ***same*** *molecular formula* because the atoms can bond to each other in different arrangements to form ***different*** *structural formulas*.

Isomers *are different compounds with the same molecular formula.*

Ethanol and dimethylether both share the molecular formula C_2H_6O and are examples of ***structural isomers*** – that is, atoms are bonded together differently.



Ethanol



Dimethylether

Silicon carbide (SiC) is made by reacting silicon dioxide (SiO_2) with powdery carbon at high temperatures. Carbon monoxide is also formed. When 100.0 kg of sand is processed, 51.4 kg SiC is recovered. What is the percent yield?

Theoretical, Actual and Percent Reaction Yields

The ***theoretical yield*** is the amount of product calculated using the stoichiometrically equivalent molar ratios obtained from the balanced equation.

In reality the theoretical yield is <u>never</u> obtained. The amount of product you obtain is called the ***actual yield***.

The ***percent yield*** (% yield) is the actual yield expressed as a percentage of the theoretical yield:

$$\% \text{ yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\%$$

If you have successive reactions, you multiply the percent yields for each reaction to determine the overall percent yield.

Silicon carbide (SiC) is made by reacting silicon dioxide (SiO_2) with powdery carbon at high temperatures. Carbon monoxide is also formed. When 100.0 kg of sand is processed, 51.4 kg SiC is recovered. What is the percent yield?

1. ***Write balanced chemical reaction.***

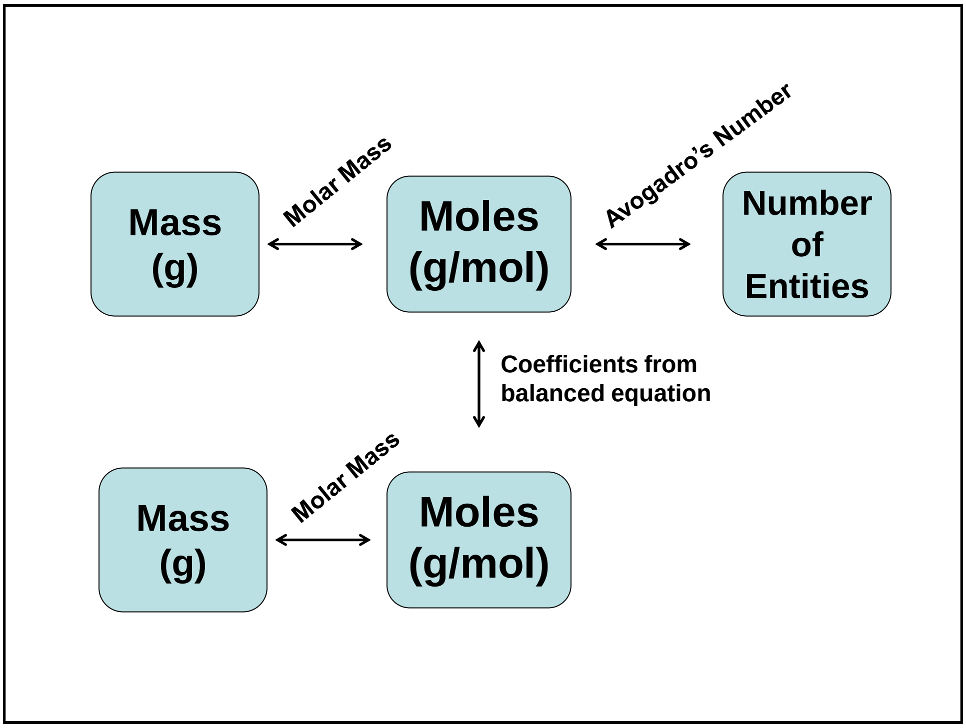
2. ***Calculate number of moles SiC produced from 100.0 kg SiO_2.***

3. ***Calculate theoretical yield based on moles SiC.***

4. ***Calculate percent yield based on 51.4 kg SiC recovered.***

Two successive reactions, D E and E F, have yields of 48% and 73% , respectively.

What is the overall percent yield for conversion of D to F?


Mass
(g)
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of
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Mass
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