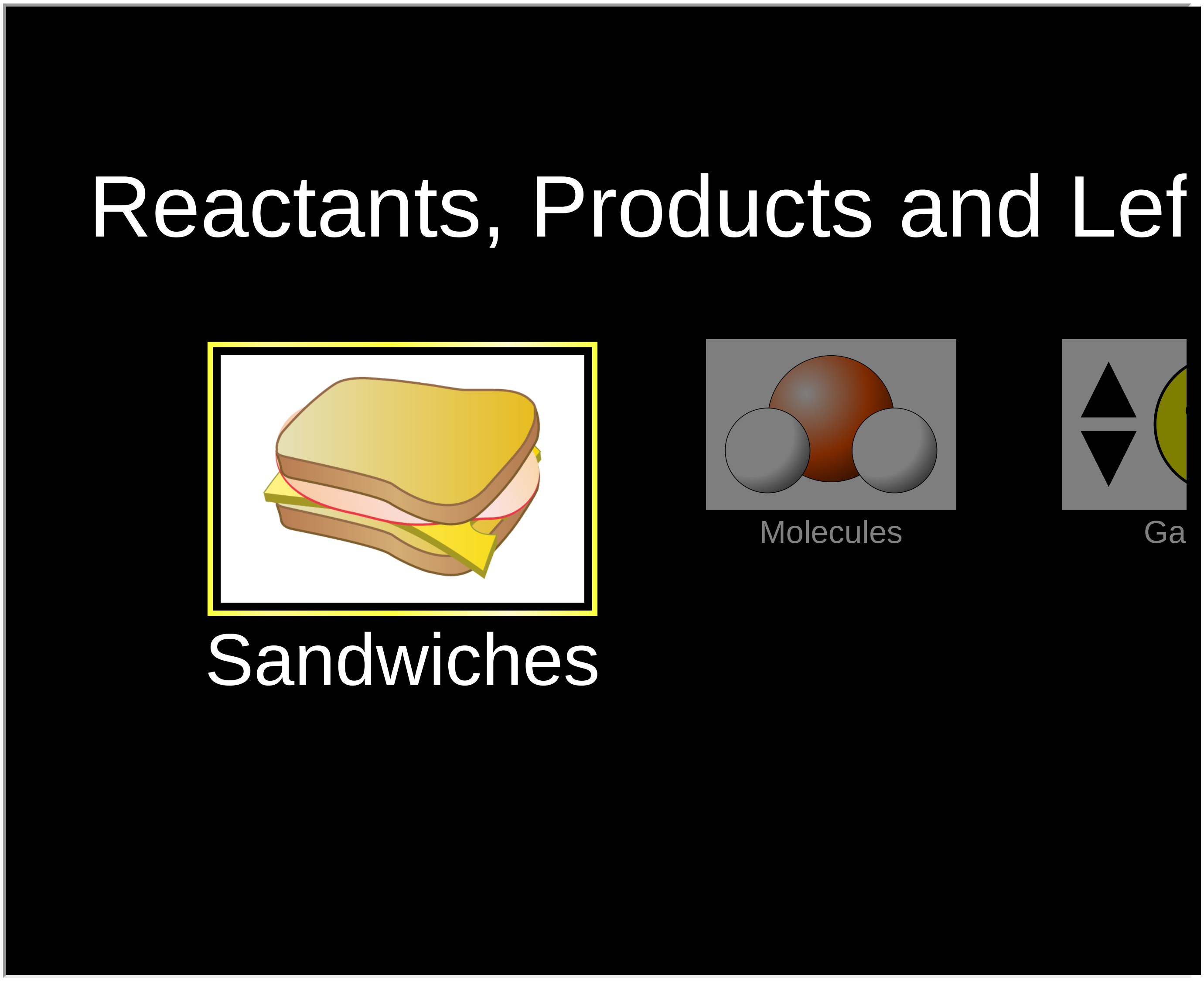
3.3: The Limiting Reagent

Example 4 from Equations and Mass Relationships(opens in new window) also illustrates the idea that one reactant in a chemical equation may be completely consumed without using up all of another. In the laboratory as well as the environment, inexpensive reagents like atmospheric O_2 are often supplied in excess. Some portion of such a reagent will be left unchanged after the reaction. Conversely, at least one reagent is usually completely consumed. When it is gone, the other excess reactants have nothing to react with and they cannot be converted to products. The substance which is used up first is the **limiting reagent**, or t**he reactant (of two or more reactants) present in an amount such that it would be completely consumed if the reaction proceeded to completion. Also called limiting reactant.**

For further understanding of limiting reagents, check out the video below or fiddle with the Phet simulation linked below the video.



Choose one of the games above to learn more about how limiting reactants work.

Example 3.3.1: Mass of Product

When 100.0 g mercury is reacted with 100.0 g bromine to form mercuric bromide, which is the limiting reagent? What mass of product will be formed?

Solution: A mixture of one or more substances dissolved in a solvent to give a homogeneous mixture. The balanced equation A representation of a chemical reaction that has values of the stoichiometric coefficients of reactants and products such that the number of atoms of each element is the same before and after the reaction.

$$\mathrm{Hg} + \mathrm{Br_2} \rightarrow \mathrm{HgBr_2}$$

tells us that according to the atomic theory, 1 mol Hg is required for each mole. That chemical amount of a substance containing the same number of units as 12 g of carbon-12. of Br_2. That is, the stoichiometric ratio S(Hg/Br_2) = 1 mol Hg/ 1 mol Br_2. Let us see how many moles of each we actually have

$$n_{\mathrm{Hg}} = 100.0\ \mathrm{g} \times \frac{1\ \mathrm{mol\ Hg}}{200.59\ \mathrm{g}} = 0.4985\ \mathrm{mol\ Hg} \tag{3.3.1}$$

$$n_{\mathrm{Br_2}} = 100.0\ \mathrm{g} \times \frac{1\ \mathrm{mol\ Br_2}}{159.80\ \mathrm{g}} = 0.6258\ \mathrm{mol\ Br_2} \tag{3.3.2}$$

When the reaction ends, 0.4985 mol Hg will have reacted with 0.4985 mol Br_2 and there will be

$$(0.6258 - 0.4985)\ \mathrm{mol\ Br_2} = 0.1273\ \mathrm{mol\ Br_2}$$

left over. Mercury is therefore the limiting reagent.

From this example you can begin to see what needs to be done to determine which of two reagents, X or Y, is limiting. We must compare the stoichiometric ratio S(X/Y) with the actual ratio of amounts of X and Y which were initially mixed together. In Example 3.3.1 this ratio of initial amounts

$$\frac{n_{\text{Hg}}(\text{initial})}{n_{\text{Br}_2}(\text{initial})}=\frac{00.4985\text{ mol Hg}}{00.6258\text{ mol Br}_2}=\frac{00.7966\text{ mol Hg}}{1\text{ mol Br}_2}$$

was less than the stoichiometric ratio

$$\text{S}\left(\frac{\text{Hg}}{\text{Br}_2}\right)=\frac{1\text{ mol Hg}}{1\text{ mol Br}_2}$$

This indicated that there was not enough Hg to react with all the bromine and mercury was the limiting reagent. The corresponding general rule, for any reagents X and Y, is

$$\text{If }\frac{n_{\text{X}}(\text{initial})}{n_{\text{Y}}(\text{initial})}\text{ is less than S}\left(\frac{\text{X}}{\text{Y}}\right)\text{, then X is limiting.}\tag{3.3.3}$$

$$\tag{3.3.4}$$

$$\text{If }\frac{n_{\text{X}}(\text{initial})}{n_{\text{Y}}(\text{initial})}\text{ is greater than S}\left(\frac{\text{X}}{\text{Y}}\right)\text{, then Y is limiting.}\tag{3.3.5}$$

These calculations can be organized as a table, with entries below the respective reactants and products in the chemical equation. We can calculate (hypothetically) how much of each reactant would be required if the other were completely consumed to demonstrate which is in excess, and which is limiting. We use the amount of limiting reagent to calculate the amount of product formed.

Calculations of respective reactants and products in chemical equation

	Hg	$+\text{Br}_2$	$\rightarrow \text{HgBr}_2$
m (g)	100	100	
M (g/mol)	200.59	159.80	360.398
n (mol)	0.4985	0.6259	--
if all Hg reacts	-0.4985	-0.4985	+0.4985
~~if all Br_2 reacts~~	~~-0.6258~~	~~-0.6258~~	
Actual Reaction Amounts	-0.4985	-0.4985	+0.4985
Actual Reaction Masses	-100	-79.66	+179.7

(Of course, when the amounts of X and Y are in exactly the stoichiometric ratio, both reagents will be completely consumed at the same time, and neither is in excess.). This general rule for determining the limiting reagent is applied in the next example.

Example 3.3.2: Limiting Reagent

Iron can be obtained by reacting the ore hematite (Fe_2O_3) with coke (C). The latter is converted to CO_2. As manager of a blast furnace you are told that you have 20.5 Mg (megagrams) of Fe_2O_3 and 2.84 Mg of coke on hand. (a) Which should you order first—another shipment of iron ore or one of coke? (b) How many megagrams of iron can you make with the materials you have?

Solution:

a) Write a balanced equation $2\text{Fe}_2\text{O}_3+3\text{C}\rightarrow 3\text{CO}_2+4\text{Fe}$

The stoichiometric ratio connecting C and Fe_2O_3 is:

$$\text{S}\left(\frac{\text{C}}{\text{Fe}_2\text{O}_3}\right)=\frac{3\text{ mol C}}{2\text{ mol Fe}_2\text{O}_3}=\frac{1.5\text{ mol C}}{1\text{ mol Fe}_2\text{O}_3}$$

The initial amounts of C and Fe_2O_3 are calculated using appropriate molar masses

$$n_{\mathrm{C}}(\text{initial}) = 2.84 \times 10^6 \mathrm{g} \times \frac{1 \text{ mol C}}{12.01 \text{ g}} = 2.36 \times 10^5 \text{mol C} \tag{3.3.6}$$

$$\tag{3.3.7}$$

$$n_{\mathrm{Fe_2O_3}}(\text{initial}) = 20.5 \times 10^6 \mathrm{g} \times \frac{1 \text{ mol Fe}_2\mathrm{O}_3}{159.69 \text{ g}} = 1.28 \times 10^5 \text{mol Fe}_2\mathrm{O}_3 \tag{3.3.8}$$

Their ratio is:

$$\frac{n_{\mathrm{C}}(\text{initial})}{n_{\mathrm{Fe_2O_3}}(\text{initial})} = \frac{2.36 \times 10^5 \text{mol C}}{1.28 \times 10^5 \text{mol Fe}_2\mathrm{O}_3} = \frac{1.84 \text{ mol C}}{1 \text{ mol Fe}_2\mathrm{O}_3}$$

Since this ratio is larger than the stoichiometric ratio, you have more than enough C to react with all the Fe_2O_3. Fe_2O_3 is the limiting reagent, and you will want to order more of it first since it will be consumed first.

b) The amount of product formed in a reaction may be calculated via an appropriate stoichiometric ratio from the amount of a reactant which was *consumed*. Some of the excess reactant C will be left over, but all the initial amount of Fe_2O_3 will be consumed. Therefore we use n_{Fe2O3} (initial) to calculate how much Fe can be obtained.

$$n_{\mathrm{Fe_2O_3}} \xrightarrow{S(\mathrm{Fe/Fe_2O_3})} n_{\mathrm{Fe}} \xrightarrow{M_{\mathrm{Fe}}} m_{\mathrm{Fe}}$$

$$m_{\mathrm{Fe}} = 1.28 \times 10^5 \text{ mol Fe}_2\mathrm{O}_3 \times \frac{4 \text{ mol Fe}}{2 \text{ mol Fe}_2\mathrm{O}_3} \times \frac{55.85 \text{ g}}{\text{mol Fe}} = 1.43 \times 10^7 \text{ g Fe}$$

This is 1.43×10^6 g, or 14.3 Mg, Fe.

As you can see from the example, in a case where there is a limiting reagent, *the initial amount of the limiting reagent must be used to calculate the amount of product formed*. Using the initial amount of a reagent present in excess would be incorrect, because such a reagent is not entirely consumed.

Liebig's law of the minimum

The concept of a limiting reagent was used by the nineteenth century German chemist Justus von Liebig (1807 to 1873) to derive an important biological and ecological law. Liebig's law of the minimum states that the essential substance available in the smallest amount relative to some critical minimum will control growth and reproduction of any species of plant or animal life. When a group, those elements that comprise a single column of the periodic table, also called family, of organisms runs out of that essential limiting reagent, the chemical reactions needed for growth and reproduction must stop. Vitamins, protein and other nutrients are essential for growth of the human body and of human populations. Similarly, the growth of algae in natural bodies of water such as Lake Erie can be inhibited by reducing the supply of nutrients such as phosphorus in the form of phosphates. It is for this reason that many states have regulated or banned the use of phosphates in detergents and are constructing treatment plants which can remove phosphates from municipal sewage before they enter lakes or streams.