6.3: Mass Calculations

The methods described in the previous section allow us to express reactants and products in terms of moles, but what if we wanted to know how many *grams* of a reactant would be required to produce a given number of *grams* of a certain product? This logical extension is, of course, trivial! In Chapter 4, we learned to express molar quantities in terms of the *masses* of reactants or products. For example, the reduction of iron (III) oxide by hydrogen gas, produces metallic iron and water. If we were to ask how many grams of elemental iron will be formed by the reduction of 1.0 grams of iron (III) oxide, we would simply use the molar stoichiometry to determine the number of moles of iron that would be produced, and then convert moles into grams using the known molar mass. For example, one gram of Fe_2O_3 can be converted into mol Fe_2O_3 by remembering that moles of a substance is equivalent to grams of that substance divided by the molar mass of that substance:

$$moles = \left(\frac{grams}{molar\ mass}\right) = \left(\frac{grams}{grams/mol}\right) = (grams) \times (mol/grams)$$

Using this approach, the mass of a reactant can be inserted into our reaction pathway as the ratio of mass-to-molar mass. This is shown here for the reduction of 1.0 gram of Fe_2O_3.

$$Given: \left(\frac{1.0 g\ Fe_2O_3}{159.70 \frac{g\ Fe_2O_3}{mol\ Fe_2O_3}}\right) \quad Find: x\ mol\ Fe$$

We set up the problem to solve for *mol product*; the general equation is:

$$(molproduct) = (molreactant) \times \left(\frac{molproduct}{molreactant}\right)$$

The stoichiometric mole ratio is set up so that mol reactant will cancel, giving a solution in *mol product*. Substituting,

$$x\ mol\ Fe = \left(\frac{1.0 g\ Fe_2O_3}{159.70 \frac{g\ Fe_2O_3}{mol\ Fe_2O_3}}\right) \times \left(\frac{2\ mol\ Fe}{1\ mol\ Fe_2O_3}\right)$$

It is often simpler to express the ratio (mass)/(molar mass) as shown below,

$$\left(\frac{1.0 g\ Fe_2O_3}{159.70 \frac{g\ Fe_2O_3}{mol\ Fe_2O_3}}\right) = (1.0 g\ Fe_2O_3) \times \left(\frac{1\ mol\ Fe_2O_3}{159.70\ g\ Fe_2O_3}\right)$$

Doing this, and rearranging,

$$x\ mol\ Fe = (1.0 g\ Fe_2O_3) \times \left(\frac{1\ mol\ Fe_2O_3}{159.70\ g\ Fe_2O_3}\right) \times \left(\frac{2\ mol\ Fe}{1\ mol\ Fe_2O_3}\right) = 0.013\ mol$$

That is, the reduction of 1.0 grams of Fe_2O_3 by excess hydrogen gas will produce 0.013 moles of elemental iron. All of these calculations are good to two significant figures based on the mass of iron (III) oxide in the original problem (1.0 grams). Note that we have *two* conversion factors (ratios) in this solution; one from mass to molar mass and the second, the stoichiometric mole ratio from the balanced chemical equation. Knowing that we have 0.013 moles of Fe, we could now convert that into grams by knowing that one mole of Fe has a mass of 55.85 grams; the yield would be 0.70 grams.

We could also modify our basic set-up so that we could find the number of *grams* of iron directly.

Here we have simply substituted the quantity (*moles molar mass*) to get mass of iron that would be produced. Again, we set up the problem to solve for *mol product*;

$$(molproduct) = (molreactant) \times \left(\frac{molproduct}{molreactant}\right)$$

In place of *mol product* and *mol reactant*, we use the expressions for *mass* and *molar mass*, as shown in the scheme above. The stoichiometric mole ratio is set up so that *mol reactant* (the ***given***) will cancel, giving a solution in *mol product*. Substituting,

$$x\ mol\ Fe\left(\frac{55.85g\ Fe}{mol\ Fe}\right)=(1.0g\ Fe_2O_3)\times\left(\frac{1\ mol\ Fe_2O_3}{159.70\ g\ Fe_2O_3}\right)\times\left(\frac{2\ mol\ Fe}{1\ mol\ Fe_2O_3}\right)$$

Rearranging and canceling units,

$$x\ g\ Fe=\left(\frac{55.85g\ Fe}{mol\ Fe}\right)\left(\frac{1\ g\ Fe_2O_3\times mol\ Fe_2O_3}{159.70\ g\ Fe_2O_3}\right)\times\left(\frac{2\ mol\ Fe}{1\ mol\ Fe_2O_3}\right)=0.70g$$

? Exercise 6.3.1

Aqueous solutions of silver nitrate and sodium chloride react in a double-replacement reaction to form a precipitate of silver chloride, according to the balanced equation shown below.

$AgNO_3$ *(aq)* + NaCl *(aq)* → AgCl *(s)* + $NaNO_3$ *(aq)*

If 3.06 grams of solid AgCl are recovered from the reaction mixture, what mass of $AgNO_3$ was present in the reactants?

? Exercise 6.3.2

Aluminum and chlorine gas react to form aluminum chloride according to the balanced equation shown in below.

2 Al *(s)* + 3 Cl_2 *(g)* → 2 $AlCl_3$ *(s)*

If 17.467 grams of chlorine gas are allowed to react with excess Al, what mass of solid aluminum chloride will be formed?

? Exercise 6.3.3

Ammonia, NH_3, is also used in cleaning solutions around the house and is produced from nitrogen and hydrogen according to the equation:

N_2 + 3 H_2 → 2 NH_3

a. If you have 6.2 moles of nitrogen what mass of ammonia could you hope to produce?
b. If you have 6.2 grams of nitrogen how many grams of hydrogen would you need?
