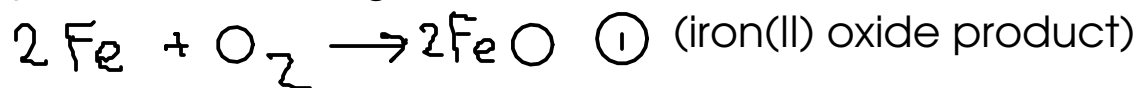


Example problem:

Determine which of the following reactions is best supported by the experimental data given.

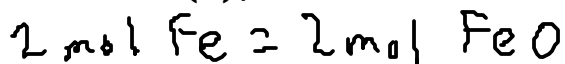


50.0 g of iron is reacted with sufficient oxygen. 71.5 g of an iron oxide are recovered

$$\text{Fe: } 55.85 \text{ g Fe} = 1 \text{ mol Fe}$$

$$50.0 \text{ g Fe} \times \frac{1 \text{ mol Fe}}{55.85 \text{ g Fe}} = 0.895255 \text{ mol Fe reacted}$$

In reaction (1), 2 moles of Fe are consumed for every 2 moles FeO formed



$$0.895255 \text{ mol Fe} \times \frac{2 \text{ mol FeO}}{2 \text{ mol Fe}} = 0.895255 \text{ mol FeO}$$

Convert to grams and compare to the actual weight of the iron oxide product collected

$$\text{FeO: } \text{Fe } 1 \times 55.85$$

$$+ \text{O } 1 \times 16.00$$

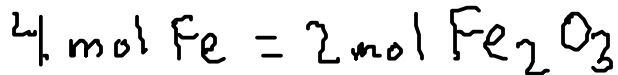
$$\hline 71.85 \text{ g FeO} = 1 \text{ mol FeO}$$

$$0.895255 \text{ mol FeO} \times \frac{71.85 \text{ g FeO}}{1 \text{ mol FeO}} = 64.3 \text{ g FeO}$$

$64.3 \text{ g} \neq 71.5 \text{ g}$  This isn't even close! FeO is not the product!

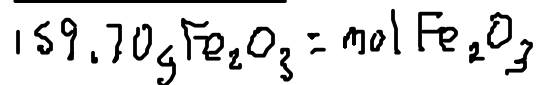
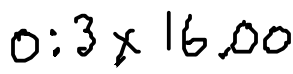
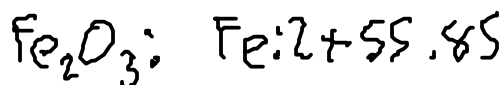
Now, check the other reaction.

In reaction (2), 4 moles of Fe are consumed for every 2 moles  $\text{Fe}_2\text{O}_3$  formed



$$0.895255 \text{ mol Fe} \times \frac{2 \text{ mol Fe}_2\text{O}_3}{4 \text{ mol Fe}} = 0.4476276 \text{ mol Fe}_2\text{O}_3$$

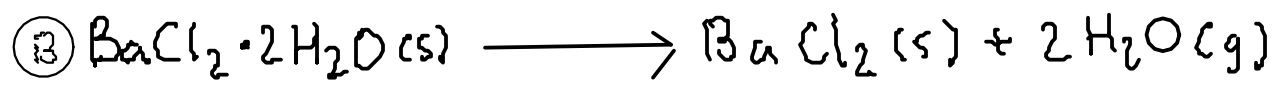
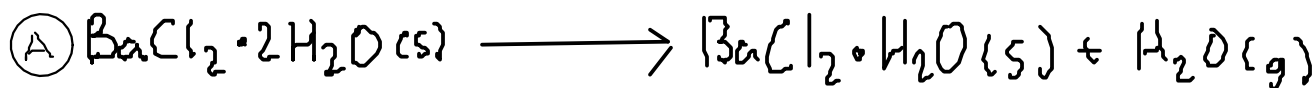
Convert to grams and compare to the actual weight of the iron oxide product collected



$$0.4476276 \text{ mol Fe}_2\text{O}_3 \times \frac{159.70 \text{ g Fe}_2\text{O}_3}{1 \text{ mol Fe}_2\text{O}_3} = \boxed{71.5 \text{ g Fe}_2\text{O}_3}$$

This value matches the amount of product observed in the experiment, so it's likely that reaction (2) is the correct one and that the iron oxide product is iron(III) oxide.

In Experiment 3B, the math is a little easier than our example problem. In all three possible reactions, the number of moles of starting material equals the number of moles of solid product:



To find out which product you have, you will need to calculate the expected weights of each solid product that could be produced in each reaction: (A), (B), and (C). As an example, let's say that you start the experiment with 2.50 g of barium chloride dihydrate: 3/5

Find out how many moles of dihydrate you have:

$$\begin{array}{l}
 \text{BaCl}_2 \cdot 2\text{H}_2\text{O} : \text{Ba} : 1 \times 137.33 \\
 \text{Cl} : 2 \times 35.45 \quad \checkmark \text{ Include the water} \\
 \text{H} : 4 \times 1.008 \quad \text{molecules!} \\
 \text{O} : 2 \times 16.00 \\
 \hline
 244.26 \text{ g BaCl}_2 \cdot 2\text{H}_2\text{O} = \text{mol BaCl}_2 \cdot 2\text{H}_2\text{O}
 \end{array}$$

$$\begin{aligned}
 2.5 \text{ g BaCl}_2 \cdot 2\text{H}_2\text{O} \times \frac{\text{mol BaCl}_2 \cdot 2\text{H}_2\text{O}}{244.26 \text{ g BaCl}_2 \cdot 2\text{H}_2\text{O}} &= \\
 &= 0.0102349 \text{ mol BaCl}_2 \cdot 2\text{H}_2\text{O}
 \end{aligned}$$

If reaction (A) is correct, one mole of barium chloride dihydrate will produce one mole of barium chloride monohydrate. In other words, the moles of barium chloride monohydrate equals the moles of barium chloride dihydrate.

$$0.0102349 \text{ mol BaCl}_2 \cdot \text{H}_2\text{O}$$

Convert moles to mass using the formula weight

$$\text{BaCl}_2 \cdot \text{H}_2\text{O} : 226.25 \text{ g BaCl}_2 \cdot \text{H}_2\text{O} = \text{mol BaCl}_2 \cdot \text{H}_2\text{O}$$

$$\begin{aligned}
 0.0102349 \text{ mol BaCl}_2 \cdot \text{H}_2\text{O} \times \frac{226.25 \text{ g BaCl}_2 \cdot \text{H}_2\text{O}}{\text{mol BaCl}_2 \cdot \text{H}_2\text{O}} &= \\
 &= 2.32 \text{ g BaCl}_2 \cdot \text{H}_2\text{O}
 \end{aligned}$$

... so if reaction (A) is the right one, we should collect about 2.32 g of solid product after the reaction is complete.

If reaction (B) is correct, one mole of barium chloride dihydrate will produce one mole of barium chloride. The moles of barium chloride equals the moles of barium chloride dihydrate.

$$0.0102349 \text{ mol BaCl}_2$$

Convert moles to mass using the formula weight

$$\text{BaCl}_2: 208.23 \text{ g BaCl}_2 = \text{mol BaCl}_2$$

$$0.0102349 \text{ mol BaCl}_2 \times \frac{208.23 \text{ g BaCl}_2}{\text{mol BaCl}_2} = \\ = 2.13 \text{ g BaCl}_2$$

... so if reaction (B) is the right one, we should collect about 2.13 g of solid product after the reaction is complete.

If reaction (C) is correct, one mole of barium chloride dihydrate will produce one mole of barium oxide..

$$0.0102349 \text{ mol BaO}$$

Convert moles to mass using the formula weight

$$\text{BaO}: 153.33 \text{ g BaO} = \text{mol BaO}$$

$$0.0102349 \text{ mol BaO} \times \frac{153.33 \text{ g BaO}}{\text{mol BaO}} = \\ = 1.57 \text{ g BaO}$$

... so if reaction (C) is the right one, we should collect about 1.57 g of solid product after the reaction is complete.

Using the chart format the lab manual uses, the results of our calculations would look like this:

This is the mass of STARTING MATERIAL (before heating!)

Substance	Formula Wt	Moles	Mass (g)
$BaCl_2 \cdot 2H_2O$	244.26	0.0102349	2.50
$BaCl_2 \cdot H_2O$	226.25	0.0102349	2.32
$BaCl_2$	208.23	0.0102349	2.13
$BaO$	153.33	0.0102349	1.57

The rest of the masses are of PRODUCTS (after heating)

Compare the calculated masses for barium chloride monohydrate, barium chloride, and barium oxide with your actual product mass.

The actual product of the reaction is the substance whose calculated mass most closely matches the actual product mass you observed during the experiment!