Thus 9.3×10^{23} silver atoms will occur in

$\frac{1}{1 \times 6.022 \times 10^{23}} \times 9.3 \times 10^{23} = 1.54$ moles of silver.

(b) We know that there are 6.022×10^{23} ammonia molecules in 1.00 mole of ammonia.
Thus there is 1 ammonia molecule present in $\frac{1}{1 \times 6.022 \times 10^{23}}$ moles of ammonia.
Thus 3.7×10^{26} ammonia molecules will occur in $\frac{1}{1 \times 6.022 \times 10^{23}} \times 3.7 \times 10^{26} = 6.1 \times 10^{2}$ moles of ammonia.

Mole calculations with equations

A balanced equation tells us the mole ratios of the reactants and products. To see this we will start by looking at the changes in mass when metals undergo combustion. You should have already seen the reaction of magnesium burning in air. It combines with oxygen to form magnesium oxide as shown in the equation.

$2Mg_{(s)} + O_{2(g)} \rightarrow 2MgO_{(s)}$

This means that:

2 atoms of magnesium react with 1 molecule of oxygen to form 2 formula units of magnesium oxide.

As we have already noted, atoms and molecules are too small for us to measure in the laboratory, so we use moles. We can read this equation as.

2 moles of magnesium react with 1 mole of oxygen to form 2 moles of magnesium oxide.

1 mol Mg = 24.31 g so that 2 mol Mg = 48.62 g

1 mol O_2 = 2 × 16.00 = 32.00 g

1 mol MgO = 24.31 + 16.00 = 40.31 so that 2 mol MgO = 80.62

From the mole ratio in the equation and these mole calculations, we can see that if we burned 48.62 g of magnesium we would need 32.00 g of oxygen and would produce 80.62 g of magnesium oxide.

No matter how much, or how little, magnesium we use, it will always combine with oxygen in the mole ratio shown, 2 mol Mg:1 mol O_2. Notice that the ratio in the equation does *not* mean that 2 g of magnesium combines with 1 g of oxygen to form 2 g of magnesium oxide. An equation tells us the ratio of numbers of particles or moles involved in a reaction, but not the masses.

If we want to work out the amount of a chemical used or produced in a chemical reaction we must:

- change all masses (weights) to moles
- write a balanced equation
- work out the mole ratio from the balanced equation
- change moles back to mass.

Here is a worked example. Follow it through and then try Set 2, Question 6.

Example

How much magnesium oxide would be produced if we burned 5.38 g magnesium?

Solution

- First we convert mass to moles. From the periodic table we know that:
  24.31 g Mg = 1 mol Mg
  Thus 1 g Mg = $\frac{1}{24.31}$ mol Mg

Thus 5.38 g Mg = 5.38 × $\frac{1}{24.31}$ = 0.221308 mol Mg

  Notice that we do not round off to 2 decimal places at this stage. Use the number in your calculator for the rest of the calculations and only round off at the final answer.
- Next we write our balanced equation to see the mole ratios used and produced when magnesium burns (combustion).
  $2Mg + O_2 \rightarrow 2MgO$
  From this equation we see that:
  2 mol Mg produces 2 mol MgO
  Thus 1 mol Mg produces 1 mol MgO.
  But we only have 0.221308 mol Mg. Using our mole ratios we see that:
  0.221308 mol Mg produces 0.221308 mol MgO
- Now we change the number of moles of MgO to grams. When we add the molar masses of the elements from the periodic table, we find the molar mass of MgO:
  1 mol MgO has a mass of 24.31 + 16 = 40.31 g
  Thus 0.221308 mol of MgO would have a mass of 40.31 × 0.221308 = 8.92 g MgO
  Notice that we have reached the end of our calculations so we can now round off.
- Finally, we have found that burning 5.38 g of Mg would produce 8.92 g MgO.
  Now try Set 2, Question 7. This question looks at the combustion of other metals. You will be able to compare mass changes in metals when they combine with oxygen.

Mole ratios are very important for chemists and chemical engineers. They use mole ratios to calculate the yield of metals expected when ores are smelted. This helps them decide whether or not it will be economic to mine the ore and extract the metal. They can also use mole calculations to ensure there is enough oxygen present in the smelting furnace and to calculate the amount of sulfur dioxide that will be emitted and have to be dealt with. Here is a sample calculation using an ore.

Example

If 20 kg of copper(I) sulfide is smelted,

(a) what will be the theoretical yield of copper?
(b) how much oxygen will be needed for complete combustion?
(c) if the actual yield of copper is 75% of the theoretical yield, how much copper will be obtained?

Solution

(a) There are three steps involved here:

- First, calculate how many moles of Cu_2S you are smelting.
  1 mole Cu_2S= 159.16 g (from the periodic table 63.55 × 2 + 32.06)
  So 20 kg $Cu_2S = \frac{m}{M} = \frac{20\ 000}{159.16}$
  = 125.6597135 moles Cu_2S
  Again, remember not to round off at this point, only when you give the final answer.
- Now use the equation to find the moles of copper produced:
  $Cu_2S_{(s)} + O_{2(g)} \rightarrow 2Cu_{(s)} + SO_{2(g)}$
  1 mole of copper sulfide + 1 mole of oxygen produces 2 moles of copper + 1 mole of sulfur dioxide
  We see that 1 mole Cu_2S will produce 2 moles Cu
  so 125.6597135 moles Cu_2S will produce 2 × 125.6597135 = 251.319427 moles Cu.
- Next change the number of moles of copper to grams:
  1 mole Cu = 63.55 g (from periodic table)
  so
  251.319427 moles Cu = 63.55 × 251.319427
  = 15 971.35 g Cu
  or 15.97 kg Cu.

  Remember you can now round off the answer.

(b) You are using 125.6597135 moles of Cu_2S.
From the equation you can see that:
1 mole of Cu_2S needs 1 mole of O_2 to be smelted,
so 125.6597135 moles of Cu_2S needs 125.6597135 moles of O_2.
Next, change moles of oxygen to grams:
1 mole O_2 = 32 grams (from periodic table)
so 125.6597135 mol = 32 × 125.6597135 g
= 4 021.11 g or 4.021 kg of oxygen used.

(c) 15.97135 kg of copper is the theoretical yield.
The actual yield is 75% of this
$\frac{75}{100} \times 15.97135 = 11.98$ kg
Now turn to Set 2 and work through Question 8 in the same way. You can use moles like this for any chemical reaction, not just those involving ores. Some other examples are given in Questions 9 and 10 of this set, and Questions 11 and 12 of Set 3.

Limiting reagents

In most of the calculations you will be asked to do, you will be told the amount of one of the chemicals reacting and you can assume that there will be plenty of the other chemical to complete the reaction. For example, if you are told that 5 g of calcium undergoes combustion to produce calcium oxide, you can assume that there is enough oxygen present for the combustion of all of the calcium.

However, if you are told how much calcium **and** how much oxygen is present, you need to check that they are in the correct mole ratio for the reaction. First you write the equation

$2Ca_{(s)} + O_{2(g)} \rightarrow 2CaO_{(s)}$

This shows us that 2 moles of Ca reacts with 1 mole of O_2 and forms 2 moles of CaO.

If you start with 2 moles of Ca and 2 moles of O_2, all the calcium will be used up before the oxygen and the reaction will stop because you run out of calcium. The **calcium will limit** your reaction.

On the other hand, if you start with 3 moles of Ca and 1 mole of O_2, all the oxygen will be used up before all the calcium has reacted. In this case the **oxygen will limit** your reaction.

The reactant that runs out first, the one that is completely used up, is called the **limiting reagent**. Whenever quantities of both reactants are known, you

must check for a limiting reagent before the problem can be solved. Here is a worked example. When you have been through this, try Set 3. Questions 13 and 14.

Example

36.46 g of hydrochloric acid is reacted with 32.69 g of zinc to produce hydrogen and a salt.

(a) Calculate the number of moles, before the reaction, of
  (i) zinc
  (ii) hydrochloric acid.
(b) Write an equation for this reaction.
(c) Which is the limiting reagent for this reaction?
(d) Calculate the number of moles of salt produced.
(e) Calculate the mass of salt produced.

Solution

(a) (i) 65.38 g Zn = 1 mol Zn (from the periodic table)

Thus 1 g Zn = $\frac{1}{65.38}$ mol Zn

and 32.69 g Zn = $\frac{32.69}{65.38}$ = 0.5 mol Zn.

(ii) Also from the periodic table, the molar mass of HCl = 1.008 + 35.45 = 36.458 g HCl.

So 1 g HCl = $\frac{1}{65.38}$ mol HCl

and 36.46 g HCl = $\frac{36.46}{36.458}$ = 1.00 mol HCl.

(b) $Zn_{(s)} + 2HCl_{(aq)} \rightarrow ZnCl_{2(aq)} + H_{2(g)}$

(c) From the equation in (b), we see that 1 mol Zn reacts with 2 mol HCl.
We have 2 mol of HCl, but only 0.5 mol Zn, so the Zn will run out first.
Thus Zn is called the **limiting reagent**.

(d) From the equation we see that:
1 mol Zn produced 1 mol $ZnCl_2$ (the salt)
thus 0.5 mol Zn will produce 0.5 mol $ZnCl_2$.

(e) We have calculated that 0.5 mol of the salt ($ZnCl_2$) will be produced. The molar mass of $ZnCl_2$ is found by:
1 mol $ZnCl_2$ = 65.38 + 2 × 35.45 = 136.28 g
Thus 0.5 mol $ZnCl_2$ has a mass of 0.5 × 136.28 = 68.14 g.

Revision Questions

Set 1

1. Calculate the amount you would measure out if you wanted exactly one mole of each of the following. Use the atomic masses shown in the periodic table in the Appendix.
   (a) iron
   (b) cobalt
   (c) chlorine gas (Cl_2)
   (d) chlorine atoms
   (e) NO_2
   (f) P_2O_5
   (g) $MgCO_3$
2. How many moles would you have in each of the following?
   (a) 78.2 g of K
   (b) 85 g of Fe
   (c) 1.91 g of Cu
   (d) 66.5 g of $AlCl_3$
   (e) 1.81×10^7 g of H_2O.
3. Calculate the mass of each of the following:
   (a) 0.5 mol of sulfur
   (b) 1.6 mol of sand (silicon dioxide)
   (c) 7.6 mol of oxygen gas
   (d) 7.6 mol of oxygen atoms
   (e) 5 mol of $Ca(OH)_2$.
4. Calculate the number of
   (a) calcium atoms in 23.4 mol of calcium
   (b) sodium ions in 0.45 mol of sodium chloride
   (c) total ions (sodium and chloride) in 0.45 mol of sodium chloride
   (d) water molecules in 21.8 mol of water
   (e) hydrogen atoms in 21.8 mol of water.
5. How many moles in each of the following?
   (a) 3.7×10^{23} bromine molecules
   (b) 3.7×10^{23} bromine atoms
   (c) 2.8×10^{30} sodium ions
   (d) 7.0×10^{15} carbon dioxide molecules
   (e) 5.6×10^{17} iron ions.

Set 2

6. 5 g lithium is converted to lithium oxide in a combustion reaction.
   (a) Write an equation for this reaction.
   (b) How many moles of lithium were burned?
   (c) How many moles of lithium oxide will be produced?
   (d) Calculate the mass of lithium oxide produced.
7. (a) Calculate the mass of:
   (i) iron(II) oxide produced by the combustion of 5 g of iron
   (ii) aluminium oxide produced by the combustion of 5 g of aluminium.
   (b) Using your answers to Questions 6 and 7(a), compare the mass changes when these lithium, iron and aluminium combine with oxygen.
8. The smelting of the copper ore chalcopyrite can be represented by the following equation:
   $2CuFeS_{2(s)} + 5O_{2(g)} \rightarrow 2Cu_{(s)} + 2FeO_{(s)} + 4SO_{2(g)}$
   (a) If 24.0 kg chalcopyrite undergoes smelting, calculate the:
   (i) moles of chalcopyrite reacting
   (ii) moles of copper formed
   (iii) mass of copper formed.
   (b) Calculate the mass actually extracted if the yield is 95%.
   (c) How many moles of oxygen gas will be needed for this reaction?
9. 3.2 g zinc reacts with excess dilute sulfuric acid, producing zinc sulfate and hydrogen gas.
   (a) Write an equation for this reaction.
   (b) Calculate the:
   (i) moles of zinc reacting
   (ii) moles of zinc sulfate produced
   (iii) mass of zinc sulfate produced.
10. Calculate the mass of
   (a) magnesium needed to react with sulfuric acid and produce 20 g of hydrogen. Include an equation in your answer.
   (b) mercury that can be obtained by heating 8.1 g of mercury(II) oxide until it all decomposes. Include an equation in your answer.

Set 3

11. When solutions of barium chloride and sodium sulfate are mixed, a white precipitate of barium sulfate is formed.
   (a) Write an equation for this reaction.
   (b) If 5.82 g of barium chloride reacts, calculate the mass of
   (i) sodium sulfate used
   (ii) barium sulfate produced.
12. Hydrogen combines with 80 g of oxygen to produce water. Calculate the:
   (a) moles of oxygen used
   (b) moles of hydrogen used
   (c) mass of hydrogen needed
   (d) mass of water produced
   (e) number of molecules of water produced.
13. (a) What is meant by the term 'limiting reagent'?
   (b) Calcium reacts with water to form calcium hydroxide and hydrogen gas. Write an equation for this reaction.
   (c) For each of the following situations identify any limiting reagent. Show all calculations.
   (i) 2 moles of calcium react with 3 moles of water.
   (ii) 1/2 mole calcium reacts with 1 mole water.
   (iii) 1 mole calcium reacts with 3 moles of water.
   (iv) 1×10^{-6} moles of calcium reacts with 2.7×10^{-6} moles of water.
14. Phosphorus undergoes combustion in the presence of oxygen to form the poisonous gas diphosphorus pentoxide.
   (a) Write an equation for this reaction.
   (b) 1.24 g of phosphorus is heated in 2.8 g of oxygen.
   (i) Identify which of these reactants is in excess.
   (ii) Identify which is the limiting reagent.
   (c) Calculate the amount of diphosphorus pentoxide formed during this reaction.
   (d) Diphosphorus pentoxide is poisonous and has been used as a nerve gas during war. What precautions are necessary if carrying out this reaction in the school laboratory?
15. (a) Define the mole and relate this to Avogadro's constant.
   (b) During a chemical reaction 3.5 mol of carbon dioxide is produced. Calculate:
   (i) the mass of carbon dioxide produced
   (ii) the number of molecules of carbon dioxide produced
   (iii) the number of atoms of oxygen present as part of the carbon dioxide.