Stoichiometric calculations.
By combining a knowledge of balancing equations with the concept of the mole, it is possible to easily calculate the masses of all reactants required or products formed in any reaction. These calculations are called STOICHIOMETRIC CALCULATIONS. How this is done is again best illustrated by some examples.

Example 1. Heating potassium chlorate, KClO₃, produces oxygen and potassium chloride. What mass of oxygen is obtained by quantitatively decomposing 3.00 g of potassium chlorate?

Start by writing a balanced formula equation for the reaction.

\[ 2\text{KClO}_3 \rightarrow 2\text{KCl} + 3\text{O}_2 \]

Then write down the relative numbers of molecules (or formula units for ionic compounds) involved as shown in the equation (the stoichiometric coefficients).

2 molecules \( \rightarrow \) 2 “molecules” + 3 molecules

If each of these numbers of molecules were multiplied by the Avogadro number, written as \( N_A \) for short, we would have

\[ 2N_A \text{molecules} \rightarrow 2N_A \text{“molecules”} + 3N_A \text{molecules} \]

As the Avogadro number of molecules of any substance is 1 mole of the substance, then the relative numbers of moles of all species in the equation would be

2 moles \( \rightarrow \) 2 moles + 3 moles

It can be seen then that the coefficients in the balanced chemical equation can be interpreted as relative numbers of moles of species as well as the relative numbers of molecules.

Usually the problem does not require calculation of the masses of all of the reactants and/or products in the reaction. More often only two species need to be involved in the calculation and they are referred to as the known (the reactant/product for which the masses are given, KClO₃ in this example) and the unknown (the reactant/product for which the masses are to be calculated, O₂ in this example).

Thus 2 moles of the known, KClO₃, produces 3 moles of the unknown, O₂.

or 1 mole of the known (KClO₃) produces 1.5 moles of the unknown (O₂)

From Topic 7, the conversion of moles to mass for any species is achieved simply by multiplying the number of moles by the molar mass (gram formula weight) of that species, as

\[ \text{amount in moles} = \frac{\text{mass in grams}}{\text{molar mass}} \]

or

\[ \text{mass in grams} = \text{amount in moles} \times \text{molar mass} \]

Amount in moles of KClO₃ in 3.00 g = \[ \frac{\text{mass KClO}_3}{\text{molar mass}} = \frac{3.00}{122.6} = 0.0245 \text{ mole} \]
Therefore, amount in moles of \( O_2 \) produced = \( 1.5 \times 0.0245 = 0.0368 \) mole
and mass of \( O_2 \) produced = molar mass of \( O_2 \times \) moles \( O_2 \\
= 32.00 \times 0.0368 = 1.18 \) g

In the above, for convenience the term “molecules” was applied to ionic species such as \( \text{NaOH} \) which do not actually exist as molecules. Strictly the more cumbersome term such as “formula units” would be more correct. Also note that in this example, the moles of unknown produced from 1 mole of the known was deduced as an additional step (shown in bold above). This equation factor can then be used directly to multiply the actual moles of the known to find the moles of the unknown. This procedure should always be followed whenever the stoichiometric coefficient of the known in the reaction equation is not already 1.

**Example 2.** What mass of carbon dioxide and of water is produced from combustion of 10.0 g of lighter fluid, butane \( (C_4H_{10}) \), and what mass of oxygen is consumed in the reaction?

Again start by writing a balanced formula equation for the reaction.

\[
2C_4H_{10} + 13O_2 \rightarrow 8CO_2 + 10H_2O
\]

Then write down the stoichiometric coefficients.

2 molecules + 13 molecules → 8 molecules + 10 molecules

2 moles + 13 moles → 8 moles + 10 moles

Using the relationship mass in grams = amount in moles × molar mass gives

\[
2 \times 58.1 \text{ g} + 13 \times 32.0 \text{ g} \rightarrow 8 \times 44.0 \text{ g} + 10 \times 18.0 \text{ g}
\]

\[
116.2 \text{ g} + 416.0 \text{ g} \rightarrow 352.0 \text{ g} + 180.0 \text{ g}
\]

Thence deduce amounts associated with the reaction of 1.0 g of butane:

1.0 g + 416.0/116.2 = 3.6 g → 352.0/116.2 = 3.0 g + 180.0/116.2 = 1.6 g

and thus amounts associated with 10.0 g of butane:

10.0 g + 36.0 g → 30.0 g + 16.0 g

\( \therefore \) 10.0 g of butane uses 36.0 g of oxygen and forms 30.0 g of carbon dioxide and 16.0 g of water.

**Example 3.** Sodium hydroxide reacts with carbon dioxide to produce sodium carbonate and water. Calculate the mass of sodium hydroxide required to prepare 53.0 g of sodium carbonate.

Using the same procedure as in Example 1,

Equation: \( 2\text{NaOH} + \text{CO}_2 \rightarrow \text{Na}_2\text{CO}_3 + \text{H}_2\text{O} \)

Ratios: 2 "molecules" 1 molecule "1 molecule" 1 molecule

therefore 2 moles 1 mole 1 mole 1 mole

from which,
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1 mole of the known, Na$_2$CO$_3$, requires 2 moles of the unknown, NaOH.

Amount of Na$_2$CO$_3$ in 53.0 g = \[
\frac{\text{mass}}{\text{molar mass}} = \frac{53.0}{(2 \times 22.99 + 12.01 + 3 \times 16.00)} = 0.500 \text{ mole}
\]

Therefore, amount of NaOH required = $2 \times 0.500 = 1.00 \text{ mole}$

and mass of NaOH = amount of NaOH $\times$ molar mass

= $1.00 \times 40.00 = 40.0 \text{ g}$

Example 4. What mass of ammonium sulfate is required to produce 3.00 mole of ammonia from the following reaction -

\[(\text{NH}_4)_2\text{SO}_4 + \text{Ca(OH)}_2 \rightarrow \text{CaSO}_4 + 2\text{NH}_3 + 2\text{H}_2\text{O}\]

Ratios: 1 mole 2 mole

whence 2 moles of the known, NH$_3$, requires 1 mole of the unknown, (NH$_4$)$_2$SO$_4$.

$\therefore$ 1 mole of the known, NH$_3$, requires 0.500 mole of the unknown, (NH$_4$)$_2$SO$_4$.

Amount of the known, NH$_3$, to be produced = 3.00 mole

Therefore amount of (NH$_4$)$_2$SO$_4$ required = $0.500 \times$ moles of NH$_3$

= $0.500 \times 3.00 = 1.50 \text{ mole}$

and mass of (NH$_4$)$_2$SO$_4$ = moles $\times$ molar mass = $1.50 \times 132.1 = 198 \text{ g}$

In this example, note that again the number of moles of the unknown required to form 1 mole of the known was deduced and then used as the factor to obtain the moles of unknown. Also, note that the ratio of the moles of only the species which are actually relevant to the calculation were set down - the moles of the other reactants/products were not required as their masses were not requested.

Example 5. “limiting reagent” problem. A piece of iron (5.59 g) is ignited in a vessel containing 1.60 g of oxygen gas to form Fe$_3$O$_4$. Deduce which reactant is in excess and calculate the amount in moles by which it is in excess.

Equation: $3\text{Fe} + 2\text{O}_2 \rightarrow \text{Fe}_3\text{O}_4$

Ratio: 3 mole 2 mole

Moles of Fe available = mass $\div$ molar mass = $5.59 \div 55.9 = 0.100 \text{ mole}$

Moles of O$_2$ available = mass $\div$ molar mass = $1.60 \div 32.0 = 0.0500 \text{ mole}$

From the ratios above, 1 mole of O$_2$ requires 1.5 mole of Fe for complete reaction.

Moles of O$_2$ available = 0.0500 which would react with $1.50 \times 0.0500 \text{ mole of Fe}$

= 0.0750 mole Fe

Amount of Fe available = 0.100 mole, therefore it is in excess by 0.025 mole
In this example, a slight variation was introduced in which a known amount of both species was given instead of just one known amount. As can be seen from the solution, essentially the same procedure applies as in the earlier examples, except in the final stage of the calculation when either of the species can be chosen as the known and the amount of the other that would be required for quantitative conversion is calculated and then compared with the actual amount available.

**Percentage yield calculations.**
In all the preceding examples, it has been assumed that complete conversion of reactants to products has occurred (quantitative conversion). The amount of product obtained when there is 100% conversion is called the **THEORETICAL YIELD**. For various reasons, conversion in reactions may not be complete or practical techniques used for separating and purifying materials may result in some loss of that material. Consequently, the amount of product obtained in practice, called the **ACTUAL YIELD**, is usually less than the theoretical yield. The usual way of expressing the yield obtained is as the **PERCENTAGE YIELD**.

The percentage yield is given by the expression

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

Note that the units used for expressing the actual and theoretical yields are usually mass units, but they can be expressed as either mass or moles, provided that the same units are used for both.

**Example 6.** Calcium (50.5 g) is heated with excess nitrogen to form calcium nitride (60.2 g). Calculate the percentage yield obtained.

Equation: \(3\text{Ca} + \text{N}_2 \rightarrow \text{Ca}_3\text{N}_2\)

Ratios: 3 moles \(\rightarrow\) 1 mole

\[\therefore 3 \text{ moles of the known, Ca, forms 1 mole of the unknown, Ca}_3\text{N}_2.\]

\[\therefore 1 \text{ mole of the known, Ca, forms } \frac{1}{3} \text{ mole of the unknown, Ca}_3\text{N}_2.\]

Amount of Ca = mass / molar mass = 50.5 / 40.08 = 1.26 mole

For 100% conversion, amount of \(\text{Ca}_3\text{N}_2\) produced = \(\frac{1}{3} \times 1.26\) mole = 0.420 mole

Mass of 0.420 mole \(\text{Ca}_3\text{N}_2\) = moles \(\times\) molar mass

\[= 0.420 \times 148.26 \text{ g} = 62.3 \text{ g (theoretical yield)}\]

Actual yield = 60.2 g

Therefore percentage yield = \(\frac{\text{actual yield}}{\text{theoretical yield}} \times 100 = \frac{60.2}{62.3} \times 100 = 96.6\%\)

**Example 7.** Cobalt(II) chloride (12.99 g) is converted to cobalt(II) sulfate (14.90 g) by a series of reactions. Calculate the percentage yield obtained.

[Insufficient information has been given to write an equation for the conversion steps, but the full equation is not needed - all that is required is the formula for the]
two compounds, cobalt(II) chloride and cobalt(II) sulfate.
1 mole of cobalt(II) chloride, CoCl₂, contains 1 mole of cobalt(II) ions.
1 mole of cobalt(II) sulfate, CoSO₄, contains 1 mole of cobalt(II) ions.
Therefore complete conversion of 1 mole of CoCl₂ to CoSO₄ would produce 1 mole of CoSO₄

Moles of CoCl₂ used = mass / molar mass = 12.99 / 129.9 = 0.100 mole
Therefore moles of CoSO₄ theoretically obtainable = 0.100 mole
Theoretical yield of CoSO₄ = moles × molar mass of CoSO₄ = 0.100 × 155.0
= 15.50 g
Actual yield = 14.90 g
Therefore % yield = \( \frac{\text{actual yield}}{\text{theoretical yield}} \times 100 = \frac{14.90 \times 100}{15.50} = 96.1 \% \)

**Objectives of this Topic.**
When you have completed this Topic, including the tutorial questions, you should have achieved the following goals:

1. Be able to calculate the mass of product obtained from a given mass of reactant, or vice versa.
2. Understand the concepts of theoretical yield, actual yield, % yield.
3. Be able to calculate the % yield for a reaction.
4. Understand the concept of limiting reagent in a reaction and apply it to calculations of amount of product obtained.

**SUMMARY**
The coefficients in a balanced equation give the relative ratios of the formula units of each reactant and product species required for complete reaction. These ratios are also the ratios of the number of moles of each participating species, a fact that could be deduced by multiplying each coefficient by the same constant, the Avogadro number. Moles of any species can be converted to mass by multiplying by the molar mass (gram formula weight) of that species. Hence the relative masses of each reactant and product can be deduced from the balanced equation and simple mole theory. Calculations of the relative masses of reactants and products is called “stoichiometry”.

If reactants are not mixed in exactly the correct ratios required to combine completely (the stoichiometric quantities), then the reactant that is consumed totally is called the limiting reagent and the excess amounts of the other reactants that are larger than the stoichiometric amounts remain unreacted.

Stoichiometric calculations allow the theoretical yield to be calculated but in practice, the actual yield obtained may well be smaller. A quantity called the percentage yield which equals the actual yield as a percentage of the theoretical yield expresses this distinction.
TUTORIAL QUESTIONS - TOPIC 9.

1. For each of the reactions represented by the following equations, supply the number of moles and masses for all of the reactants and products, assuming complete reaction in which 1.0 mole of the specified reactant is used or 1.0 mole of the specified product is produced. Quote your answers to the nearest gram of reactants and products. Part (a) is an example:

(a) For 1.0 mole of SO₂ reacting,

Equation: \( \text{O}_2(g) + 2\text{SO}_2(g) \rightarrow 2\text{SO}_3(g) \)

- Mole ratios of reactants/products: \( \frac{1}{2} \frac{2}{2} \) [from the equation coefficients]
- Moles required: 0.50 mole \( \textbf{1.0 mole} \) 1.0 mole [from the question data and the above ratios]
- Molar masses: 32.0 g mol\(^{-1}\) 64.1 g mol\(^{-1}\) 80.1 g mol\(^{-1}\) [from atomic weight data]
- Masses required: \( 0.50 \times 32.0 \text{ g} = 16 \text{ g} \)
\( 1.0 \times 64.1 \text{ g} = 64 \text{ g} \)
\( 1.0 \times 80.1 \text{ g} = 80 \text{ g} \)

(b) For 1.0 mole of CaCO₃ reacting,

Equation: \( \text{CaCO}_3 \rightarrow \text{CaO} + \text{CO}_2 \)

- Mole ratios: \( \frac{1}{1} \frac{1}{1} \)
- Moles required: \( \textbf{1.0 mole} \)
- Molar masses: \( \text{Masses required:} \)

(c) For 1.0 mole of Fe reacting

Equation: \( 2\text{Fe} + 3\text{Cl}_2 \rightarrow 2\text{FeCl}_3 \)

- Mole ratios: \( \frac{1}{1} \frac{3}{2} \)
- Moles required: \( \textbf{1.0 mole} \)
- Molar masses: \( \text{Masses required:} \)

(d) For 1.0 mole of H₂O₂ reacting,

Equation: \( 2\text{H}_2\text{O}_2 \rightarrow 2\text{H}_2\text{O} + \text{O}_2 \)

- Mole ratios: \( \frac{1}{2} \frac{2}{1} \)
- Moles required: \( \textbf{1.0 mole} \)
- Molar masses: \( \text{Masses required:} \)

(e) For 1.0 mole of O₂ produced,

Equation: \( 2\text{Pb(NO}_3)_2 \rightarrow 2\text{PbO} + \text{O}_2 + 4\text{NO}_2 \)

- Mole ratios: \( \frac{1}{2} \frac{1}{1} \frac{4}{1} \)
- Moles required: \( \textbf{1.0 mole} \)
- Molar masses: \( \text{Masses required:} \)
(f) For 1.0 mole of N₂ reacting,

Equation: \( \text{N}_2 + 3\text{H}_2 \rightarrow 2\text{NH}_3 \)
Mole ratios:
Moles required: \( 1.0 \text{ mole} \)
Molar masses:
Masses required:

(g) For 1.0 mole of H₂ reacting,

Equation: \( \text{N}_2 + 3\text{H}_2 \rightarrow 2\text{NH}_3 \)
Mole ratios:
Moles required: \( 1.0 \text{ mole} \)
Molar masses:
Masses required:

(h) For 1.0 mole of NH₃ produced,

Equation: \( \text{N}_2 + 3\text{H}_2 \rightarrow 2\text{NH}_3 \)
Mole ratios:
Moles required: \( 1.0 \text{ mole} \)
Molar masses:
Masses required:

(i) For 1.0 mole of Pb(NO₃)₂ reacting

Equation: \( 2\text{Pb(NO}_3)_2 \rightarrow 2\text{PbO} + \text{O}_2 + 4\text{NO}_2 \)
Mole ratios:
Moles required: \( 1.0 \text{ mole} \)
Molar masses:
Masses required:

(j) For 1.0 mole of ZnS reacting,

Equation: \( 2\text{ZnS} + 3\text{O}_2 \rightarrow 2\text{ZnO} + 2\text{SO}_2 \)
Mole ratios:
Moles required: \( 1.0 \text{ mole} \)
Molar masses:
Masses required:

(k) For 1.0 mole of O₂ reacting,

Equation: \( 2\text{ZnS} + 3\text{O}_2 \rightarrow 2\text{ZnO} + 2\text{SO}_2 \)
Mole ratios:
Moles required: \( 1.0 \text{ mole} \)
Molar masses:
Masses required:
(l) For 1.0 mole of NaHCO₃ reacting,

Equation: \[ 2 \text{NaHCO}_3 \rightarrow \text{Na}_2\text{CO}_3 + \text{CO}_2 + \text{H}_2\text{O} \]
Mole ratios:
Moles required: 1.0 mole
Molar masses:
Masses required:

(m) For 1.0 mole of C₂H₆ reacting,

Equation: \[ 2 \text{C}_2\text{H}_6 + 7\text{O}_2 \rightarrow 4\text{CO}_2 + 6\text{H}_2\text{O} \]
Mole ratios:
Moles required: 1.0 mole
Molar masses:
Masses required:

(n) For 1.0 mole of C₆H₁₄ reacting,

Equation: \[ 2 \text{C}_6\text{H}_{14} + 19\text{O}_2 \rightarrow 12\text{CO}_2 + 14\text{H}_2\text{O} \]
Mole ratios:
Moles required: 1.0 mole
Molar masses:
Masses required:

(o) For 1.0 mole of CO₂ produced,

Equation: \[ 2 \text{C}_6\text{H}_{14} + 19\text{O}_2 \rightarrow 12\text{CO}_2 + 14\text{H}_2\text{O} \]
Mole ratios:
Moles required: 1.0 mole
Molar masses:
Masses required:

(p) For 1.0 mole of Li₃N produced,

Equation: \[ 6\text{Li} + \text{N}_2 \rightarrow 2\text{Li}_3\text{N} \]
Mole ratios:
Moles required: 1.0 mole
Molar masses:
Masses required:

(q) For 1.0 mole of Cl₂ produced,

Equation: \[ 4\text{HCl} + \text{O}_2 \rightarrow 2\text{H}_2\text{O} + 2\text{Cl}_2 \]
Mole ratios:
Moles required: 1.0 mole
Molar masses:
Masses required:
2. Hydrogen gas combines with chlorine gas to form hydrogen chloride. If 2.00 mole of hydrogen reacts with excess chlorine, what mass of hydrogen chloride can be obtained?

3. What mass of carbon dioxide is produced when 18 g of carbon is burnt in excess oxygen?

4. Hydrogen and chlorine gases combine to form hydrogen chloride gas, HCl. Calculate the minimum mass of hydrogen gas and of chlorine gas needed to produce 10.0 g of hydrogen chloride.

5. Calculate the mass of chlorine needed to produce 5.85 g of sodium chloride when excess sodium is heated with chlorine gas.

6. Iron burns in air to form the compound Fe$_3$O$_4$.
   (a) Write the equation for the reaction.
   (b) How many moles of oxygen gas are needed to burn 1.0 mole of iron?
   (c) What is the mass of this amount of oxygen gas?
   (d) Would 0.050 mole of oxygen be enough to cause 5.6 g of iron to react completely to form Fe$_3$O$_4$?

7. The gas N$_2$O can be prepared by heating a mixture of ammonium chloride and sodium nitrate according to the following equation:

   \[
   \text{NH}_4\text{Cl} + \text{NaNO}_3 \rightarrow \text{N}_2\text{O} + \text{NaCl} + 2\text{H}_2\text{O}
   \]

   (i) What mass of N$_2$O can be prepared from 100.0 g of ammonium chloride?

   (ii) The mass of N$_2$O actually obtained from this experiment was 75.4 g. Calculate the percentage yield.

8. The oxide of iron, Fe$_3$O$_4$, can be converted to iron metal by heating with excess hydrogen gas. What mass of Fe$_3$O$_4$ is needed to produce 112 g of iron?

9. Sodium nitrate decomposes on heating to form sodium nitrite and oxygen gas.
   (i) What mass of sodium nitrate would be required to produce 64.0 g of oxygen gas if 100 % conversion occurred?

   (ii) If the actual yield from this reaction was only 80.0 %, what mass of sodium nitrate would have to be used in order to produce 64.0 g of oxygen?

10. Carbon (1.20 g) is burned in a closed container with 8.00 g of oxygen.
    (i) What mass of carbon dioxide would be formed?

    (ii) Which reagent is in excess and what mass of it remains after the reaction?
11. Aluminium sulfide, Al\textsubscript{2}S\textsubscript{3}, can be prepared by reacting aluminium with sulfur.

(i) What mass of aluminium sulfide can be prepared from 10.0 g of aluminium mixed with 15.0 g of sulfur?

(ii) Which is the non-limiting reagent and by what mass is it in excess?

12. Sulfur dioxide gas can be obtained by heating solid sodium hydrogensulfite with sulfuric acid as shown in the following equation:

\[
2\text{NaHSO}_3 + \text{H}_2\text{SO}_4 \rightarrow 2\text{SO}_2 + 2\text{H}_2\text{O} + \text{Na}_2\text{SO}_4
\]

(i) Industrially, instead of using moles as gram formula weights for the molar masses of reactants and products, often units of tonne formula weights are preferred. How many tonnes of sulfur dioxide gas can be obtained from 1.00 tonne of sodium hydrogensulfite?

(ii) The standard glob was a unit of mass used by ancient Martians before they succumbed to the loss of atmosphere through the greenhouse effect. How many globs of SO\textsubscript{2} can be prepared from 5.0 globs of NaHSO\textsubscript{3} in the reaction represented by the above equation?

13. Calculate the mass of sodium hydroxide required to react exactly with 0.50 mole of nitric acid, HNO\textsubscript{3}.

14. Nickel(II) nitrate-6-water has the formula Ni(NO\textsubscript{3})\textsubscript{2}.6H\textsubscript{2}O. If 29.1 g of this salt were dissolved completely in water, calculate the following for the resulting solution:

(a) The number of moles of Ni\textsuperscript{2+} ions present.
(b) The number of moles of NO\textsuperscript{3-} ions present.
(c) The number of nickel ions present.
(d) The number of nitrate ions present.

15. A piece of pure zinc weighing 15.8 g is dissolved completely in hydrochloric acid. Calculate:

(a) the mass of hydrogen gas formed
(b) the mass of zinc chloride that could be formed by evaporating the water from the resultant solution.
1. (b) For 1.0 mole of CaCO₃ reacting,

   Equation: \( \text{CaCO}_3 \rightarrow \text{CaO} + \text{CO}_2 \)

   Mole ratios: \(1 : 1 : 1\)
   Moles required: \(1.0 : 1.0 : 1.0\)
   Molar masses: \(100.1 \text{ g mol}^{-1} : 56.1 \text{ g mol}^{-1} : 44.0 \text{ g mol}^{-1}\)
   Masses required: \(1.0 \times 100.1 \) \(= 100 \text{ g}\)
   \(1.0 \times 56.1 \) \(= 56 \text{ g}\)
   \(1.0 \times 44.0 \) \(= 44 \text{ g}\)

(c) For 1.0 mole of Fe reacting

   Equation: \(2\text{Fe} + 3\text{Cl}_2 \rightarrow 2\text{FeCl}_3\)

   Mole ratios: \(2 : 3 : 2\)
   Moles required: \(1.0 : 1.5 : 1.0\)
   Molar masses: \(55.8 \text{ g mol}^{-1} : 70.9 \text{ g mol}^{-1} : 162.2 \text{ g mol}^{-1}\)
   Masses required: \(1.0 \times 55.8 \) \(= 56 \text{ g}\)
   \(1.5 \times 70.9 \) \(= 106 \text{ g}\)
   \(1.0 \times 162.2 \) \(= 162 \text{ g}\)

(d) For 1.0 mole of \( \text{H}_2\text{O}_2 \) reacting,

   Equation: \(2\text{H}_2\text{O}_2 \rightarrow 2\text{H}_2\text{O} + \text{O}_2\)

   Mole ratios: \(2 : 2 : 1\)
   Moles required: \(1.0 : 1.0 : 0.50\)
   Molar masses: \(34.0 \text{ g mol}^{-1} : 18.0 \text{ g mol}^{-1} : 32.0 \text{ g mol}^{-1}\)
   Masses required: \(1.0 \times 34.0 \) \(= 34 \text{ g}\)
   \(1.0 \times 18.0 \) \(= 18 \text{ g}\)
   \(0.50 \times 32.0 \) \(= 16 \text{ g}\)

(e) For 1.0 mole of \( \text{O}_2 \) produced,

   Equation: \(2\text{Pb(NO}_3)_2 \rightarrow 2\text{PbO} + \text{O}_2 + 4\text{NO}_2\)

   Mole ratios: \(2 : 2 : 1 : 4\)
   Moles required: \(2.0 : 2.0 : 1.0 : 4.0\)
   Molar masses: \(331.2 \text{ g mol}^{-1} : 223.2 \text{ g mol}^{-1} : 32.0 \text{ g mol}^{-1} : 46.0 \text{ g mol}^{-1}\)
   Masses required: \(2.0 \times 331.2 \) \(= 662 \text{ g}\)
   \(2.0 \times 223.2 \) \(= 446 \text{ g}\)
   \(1.0 \times 32.0 \) \(= 32 \text{ g}\)
   \(4.0 \times 46.0 \) \(= 184 \text{ g}\)
(f) For 1.0 mole of N\textsubscript{2} reacting,

\begin{align*}
\text{Equation: } & \quad \text{N}_2 + 3\text{H}_2 \rightarrow 2\text{NH}_3 \\
\text{Mole ratios: } & \quad 1 \quad 3 \quad 2 \\
\text{Moles required: } & \quad 1.0 \quad 3.0 \quad 2.0 \\
\text{Molar masses: } & \quad 28.0 \text{ g mol}^{-1} \quad 2.0 \text{ g mol}^{-1} \quad 17.0 \text{ g mol}^{-1} \\
\text{Masses required: } & \quad 1.0 \times 28.0 = 28 \text{ g} \quad 3.0 \times 2.0 = 6 \text{ g} \quad 2.0 \times 17.0 = 34 \text{ g}
\end{align*}

(g) For 1.0 mole of H\textsubscript{2} reacting,

\begin{align*}
\text{Equation: } & \quad \text{N}_2 + 3\text{H}_2 \rightarrow 2\text{NH}_3 \\
\text{Mole ratios: } & \quad 1 \quad 3 \quad 2 \\
\text{Moles required: } & \quad 0.333 \quad 1.0 \quad 0.667 \\
\text{Molar masses: } & \quad 28.0 \text{ g mol}^{-1} \quad 2.0 \text{ g mol}^{-1} \quad 17.0 \text{ g mol}^{-1} \\
\text{Masses required: } & \quad 0.333 \times 28 = 9.3 \text{ g} \quad 1.0 \times 2.0 = 2.0 \text{ g} \quad 0.667 \times 17.0 = 11.3 \text{ g}
\end{align*}

(h) For 1.0 mole of NH\textsubscript{3} produced,

\begin{align*}
\text{Equation: } & \quad \text{N}_2 + 3\text{H}_2 \rightarrow 2\text{NH}_3 \\
\text{Mole ratios: } & \quad 1 \quad 3 \quad 2 \\
\text{Moles required: } & \quad 0.50 \quad 1.5 \quad 1.0 \\
\text{Molar masses: } & \quad 28.0 \text{ g mol}^{-1} \quad 2.0 \text{ g mol}^{-1} \quad 17.0 \text{ g mol}^{-1} \\
\text{Masses required: } & \quad 0.50 \times 28.0 = 14 \text{ g} \quad 1.5 \times 2.0 = 3 \text{ g} \quad 1.0 \times 17.0 = 17 \text{ g}
\end{align*}

(i) For 1.0 mole of Pb(NO\textsubscript{3})\textsubscript{2} reacting

\begin{align*}
\text{Equation: } & \quad 2\text{Pb(NO}_3\text{)}_2 \rightarrow 2\text{PbO} + \text{O}_2 + 4\text{NO}_2 \\
\text{Mole ratios: } & \quad 2 \quad 2 \quad 1 \quad 4 \\
\text{Moles required: } & \quad 1.0 \quad 1.0 \quad 0.50 \quad 2.0 \\
\text{Molar masses: } & \quad 331 \text{ g mol}^{-1} \quad 223 \text{ g mol}^{-1} \quad 32 \text{ g mol}^{-1} \quad 46 \text{ g mol}^{-1} \\
\text{Masses required: } & \quad 1.0 \times 331 = 331 \text{ g} \quad 1.0 \times 223 = 223 \text{ g} \quad 0.50 \times 32 = 16 \text{ g} \quad 2.0 \times 46 = 92 \text{ g}
\end{align*}
(j) For 1.0 mole of ZnS reacting,

Equation: \( 2\text{ZnS} + 3\text{O}_2 \rightarrow 2\text{ZnO} + 2\text{SO}_2 \)

Mole ratios: \( 2 \) \( 3 \) \( 2 \) \( 2 \)

Moles required: \( 1.0 \)

Molar masses: \( 97.5 \text{ g mol}^{-1} \) \( 32.0 \text{ g mol}^{-1} \) \( 81.4 \text{ g mol}^{-1} \) \( 64.1 \text{ g mol}^{-1} \)

Masses required: \( 1.0 \times 97.5 \text{ g} \) \( 1.5 \times 32.0 \text{ g} \) \( 1.0 \times 81.4 \text{ g} \) \( 1.0 \times 64.1 \text{ g} \)

\( = 97.5 \text{ g} \) \( = 48.0 \text{ g} \) \( = 81.4 \text{ g} \) \( = 64.1 \text{ g} \)

(k) For 1.0 mole of \( \text{O}_2 \) reacting,

Equation: \( 2\text{ZnS} + 3\text{O}_2 \rightarrow 2\text{ZnO} + 2\text{SO}_2 \)

Mole ratios: \( 2 \) \( 3 \) \( 2 \) \( 2 \)

Moles required: \( 2/3 = 0.667 \)

Molar masses: \( 97.5 \text{ g mol}^{-1} \) \( 32.0 \text{ g mol}^{-1} \) \( 81.4 \text{ g mol}^{-1} \) \( 64.1 \text{ g mol}^{-1} \)

Masses required: \( 0.667 \times 97.5 \text{ g} \) \( 1.0 \times 32.0 \text{ g} \) \( 0.667 \times 81.4 \text{ g} \) \( 0.667 \times 64.1 \text{ g} \)

\( = 65.0 \text{ g} \) \( = 32.0 \text{ g} \) \( = 54.3 \text{ g} \) \( = 42.7 \text{ g} \)

(l) For 1.0 mole of \( \text{NaHCO}_3 \) reacting,

Equation: \( 2\text{NaHCO}_3 \rightarrow \text{Na}_2\text{CO}_3 + \text{CO}_2 + \text{H}_2\text{O} \)

Mole ratios: \( 2 \) \( 1 \) \( 1 \) \( 1 \)

Moles required: \( 1.0 \)

Molar masses: \( 84.0 \text{ g mol}^{-1} \) \( 106.0 \text{ g mol}^{-1} \) \( 44.0 \text{ g mol}^{-1} \) \( 18.0 \text{ g mol}^{-1} \)

Masses required: \( 1.0 \times 84.0 \text{ g} \) \( 0.50 \times 106.0 \text{ g} \) \( 0.50 \times 44.0 \text{ g} \) \( 0.50 \times 18.0 \text{ g} \)

\( = 84 \text{ g} \) \( = 53 \text{ g} \) \( = 22 \text{ g} \) \( = 9 \text{ g} \)

(m) For 1.0 mole of \( \text{C}_2\text{H}_6 \) reacting,

Equation: \( 2\text{C}_2\text{H}_6 + 7\text{O}_2 \rightarrow 4\text{CO}_2 + 6\text{H}_2\text{O} \)

Mole ratios: \( 2 \) \( 7 \) \( 4 \) \( 6 \)

Moles required: \( 1.0 \)

Molar masses: \( 30.1 \text{ g mol}^{-1} \) \( 32.0 \text{ g mol}^{-1} \) \( 44.0 \text{ g mol}^{-1} \) \( 18.0 \text{ g mol}^{-1} \)

Masses required: \( 1.0 \times 30.1 \text{ g} \) \( 3.5 \times 32.0 \text{ g} \) \( 2.0 \times 44.0 \text{ g} \) \( 3.0 \times 18.0 \text{ g} \)

\( = 30 \text{ g} \) \( = 112 \text{ g} \) \( = 88 \text{ g} \) \( = 54 \text{ g} \)
(n) For 1.0 mole of C_{6}H_{14} reacting,

Equation: \[ \text{2C}_{6}\text{H}_{14} + 19\text{O}_{2} \rightarrow 12\text{CO}_{2} + 14\text{H}_{2}\text{O} \]

Mole ratios:
\[
\begin{array}{cccc}
2 & 19 & 12 & 14 \\
\end{array}
\]

Moles required:
\[
\begin{array}{cccc}
1.0 & 9.5 & 6.0 & 7.0 \\
\end{array}
\]

Molar masses:
\[
\begin{array}{cccc}
86.2 \text{ g mol}^{-1} & 32.0 \text{ g mol}^{-1} & 44.0 \text{ g mol}^{-1} & 18.0 \text{ g mol}^{-1} \\
\end{array}
\]

Masses required:
\[
\begin{array}{cccc}
86 \text{ g} & 304 \text{ g} & 264 \text{ g} & 126 \text{ g} \\
\end{array}
\]

(o) For 1.0 mole of CO_{2} produced,

Equation: \[ \text{2C}_{6}\text{H}_{14} + 19\text{O}_{2} \rightarrow 12\text{CO}_{2} + 14\text{H}_{2}\text{O} \]

Mole ratios:
\[
\begin{array}{cccc}
2 & 19 & 12 & 14 \\
\end{array}
\]

Moles required:
\[
\begin{array}{cccc}
2/12 = 0.167 & 19/12 = 1.58 & 1.0 & 14/12 = 1.17 \\
\end{array}
\]

Molar masses:
\[
\begin{array}{cccc}
86.2 \text{ g mol}^{-1} & 32.0 \text{ g mol}^{-1} & 44.0 \text{ g mol}^{-1} & 18.0 \text{ g mol}^{-1} \\
\end{array}
\]

Masses required:
\[
\begin{array}{cccc}
14.4 \text{ g} & 50.6 \text{ g} & 44 \text{ g} & 21 \text{ g} \\
\end{array}
\]

(p) For 1.0 mole of Li_{3}N produced,

Equation: \[ 6\text{Li} + \text{N}_{2} \rightarrow 2\text{Li}_{3}\text{N} \]

Mole ratios:
\[
\begin{array}{cccc}
6 & 1 & 2 \\
\end{array}
\]

Moles required:
\[
\begin{array}{cccc}
3.0 & 0.50 & 1.0 \\
\end{array}
\]

Molar masses:
\[
\begin{array}{cccc}
6.9 \text{ g mol}^{-1} & 28.0 \text{ g mol}^{-1} & 34.8 \text{ g mol}^{-1} \\
\end{array}
\]

Masses required:
\[
\begin{array}{cccc}
21 \text{ g} & 14 \text{ g} & 35 \text{ g} \\
\end{array}
\]

(q) For 1.0 mole of Cl_{2} produced,

Equation: \[ 4\text{HCl} + \text{O}_{2} \rightarrow 2\text{H}_{2}\text{O} + 2\text{Cl}_{2} \]

Mole ratios:
\[
\begin{array}{cccc}
4 & 1 & 2 & 2 \\
\end{array}
\]

Moles required:
\[
\begin{array}{cccc}
2.0 & 0.50 & 1.0 & 1.0 \\
\end{array}
\]

Molar masses:
\[
\begin{array}{cccc}
36.5 \text{ g mol}^{-1} & 32.0 \text{ g mol}^{-1} & 18.0 \text{ g mol}^{-1} & 70.9 \text{ g mol}^{-1} \\
\end{array}
\]

Masses required:
\[
\begin{array}{cccc}
73 \text{ g} & 16 \text{ g} & 18 \text{ g} & 71 \text{ g} \\
\end{array}
\]
Worked solutions to questions 2 - 15 are provided on pages 15 - 20.

2. 146 g
3. 66 g
4. 0.277 g H₂ and 9.71 g Cl₂
5. 3.55 g
6. (a) 3Fe + 2O₂ → Fe₃O₄
   (b) 0.67 mol
   (c) 21 g
   (d) No. 5.6 g Fe (0.10 mol) requires 0.067 mol O₂.
7. (i) 82.30 g (ii) 91.6 %
8. 154 g
9. (i) 340 g (ii) 425 g
10. (i) 4.40 g (ii) O₂ by 4.80 g
11. (i) 23.4 g (ii) Al by 1.59 g
12. (i) 0.615 ton (ii) 3.1 globs
13. 20 g
14. (a) 0.100 mol (b) 0.200 mol (c) 6.02 × 10²² (d) 12.0 × 10²²
15. (a) 0.488 g (b) 33.0 g

WORKED SOLUTIONS FOR QUESTIONS 2 - 15.

2. Balanced equation: H₂ + Cl₂ → 2HCl
Mole ratios: 1 mol 2 mol
Moles of H₂ available = 2.00 mol
∴ moles of HCl produced = 2 × 2.00 mol = 4.00 mol
Mass of HCl = moles of HCl × molar mass of HCl
= 4.00 × (1.008 + 35.45) = 146 g
3. Balanced equation: \( C + O_2 \rightarrow CO_2 \)
Mole ratios: \[ 1 \text{ mol} : 1 \text{ mol} \]
Mass of carbon available = 18 g
\[ \therefore \text{ moles of C available} = \frac{18}{12.01} = 1.50 \text{ mol} \]
From the mole ratios above, moles of CO\(_2\) formed = 1.50 mol
Mass of carbon dioxide = moles × molar mass of CO\(_2\)
\[ = 1.50 \times (12.01 + 2 \times 16.00) = 66 \text{ g} \]

4. Balanced equation: \( H_2 + Cl_2 \rightarrow 2HCl \)
Mole ratios: \[ 1 \text{ mol} : 1 \text{ mol} : 2 \text{ mol} \]
Mass of hydrogen chloride = 10.0 g
\[ \therefore \text{ moles of HCl} = \frac{10.0}{(1.01 + 35.45)} = 0.274 \text{ mol} \]
1 mole of HCl requires \( \frac{1}{2} \) mole of H\(_2\) and \( \frac{1}{2} \) mole of Cl\(_2\)
\[ \therefore \text{ moles of H}_2 = \text{ moles of Cl}_2 = 0.500 \times 0.274 \text{ mol} = 0.137 \text{ mol} \]
Mass of hydrogen = moles × molar mass of H\(_2\) = 0.137 × 2.02 g = 0.277 g
Mass of chlorine = moles × molar mass of Cl\(_2\) = 0.137 × 70.90 g = 9.71 g

5. Balanced Equation: \( 2Na + Cl_2 \rightarrow 2NaCl \)
Mole ratios: \[ 1 \text{ mol} : 2 \text{ mol} \]
Mass of sodium chloride = 5.85 g
\[ \therefore \text{ moles of NaCl} = \frac{5.85}{58.5} = 0.100 \text{ mol} \]
Moles of Cl\(_2\) required = \( \frac{1}{2} \times \) moles of NaCl = 0.500 × 0.100 = 0.0500 mol
Mass of Cl\(_2\) = moles × molar mass of Cl\(_2\) = 0.0500 × 70.90 g = 3.55 g

6. (a) \( 3Fe + 2O_2 \rightarrow Fe_3O_4 \)

(b) From the stoichiometric coefficients in the equation, 3 moles of Fe require 2 moles of O\(_2\) so 1.0 mole of Fe requires 2/3 moles of O\(_2\) = 0.667 mol

(c) Molar mass of O\(_2\) = 2 × 16.00 g = 32.00 g \text{ mol}^{-1}
\[ \therefore \text{ 0.667 mole of O}_2 \text{ has a mass} = 0.667 \times 32.00 = 21 \text{ g} \]

(d) From the equation, 1 mole of O\(_2\) requires 1.5 moles of Fe for complete reaction.
\[ \therefore \text{ 0.050 mole of O}_2 \text{ would require} = 0.050 \times 1.5 = 0.075 \text{ mole of Fe.} \]
Mass of 0.075 mole of Fe = 0.075 × molar mass of Fe = 0.075 × 55.85 g = 4.2 g.
Thus there would be an excess of iron = 5.6 – 4.2 = 1.4 g
**Alternative method:**

Moles of Fe = mass / molar mass Fe = 5.6 / 55.85 = 0.10 mol

From the equation, 1 mole of Fe requires 2/3 mole of O₂ = 0.667 mol

∴ moles of O₂ needed to react with 0.10 mole of Fe = 0.0667 mole, so 0.050 mole of O₂ is insufficient.

7. (i) From the given equation,

1 mole NH₄Cl reacting leads to the formation of 1 mole N₂O.

Moles of NH₄Cl = mass / molar mass = 100.0 / 53.49 = 1.870 mol

∴ moles of N₂O produced = 1.870 mol

and mass of N₂O = moles × molar mass = 1.870 × 44.01 g = 82.30 g

(ii) Theoretical yield = 82.30 g

Actual yield = 75.4 g

% yield = \( \frac{\text{actual yield}}{\text{theoretical yield}} \times 100 \) = \( \frac{75.4}{82.3} \times 100 \) = 91.6 %

8. Equation: Fe₃O₄ + 4H₂ → 3Fe + 4H₂O

Mole ratios: 1 mol 3 mol

∴ 1 mole of Fe requires 1/3 mole of Fe₃O₄ = 0.333 mol

Moles of Fe = mass / molar mass of Fe = 112 / 55.85 = 2.00 mol

Moles of Fe₃O₄ = 0.333 × 2.00 = 0.667 mol

Mass of Fe₃O₄ = moles × molar mass of Fe₃O₄ = 0.667 × 231.5 = 154 g

9. (i) Equation: 2NaNO₃ → 2NaNO₂ + O₂

Mole ratios: 2 mol 1 mol

1 mole of O₂ requires 2 moles of NaNO₃

Moles of O₂ = mass / molar mass of O₂ = 64.0 / 32.0 = 2.00 mol

∴ moles of NaNO₃ = 2 × 2.00 = 4.00 mol

and mass of sodium nitrate = moles × molar mass of NaNO₃ = 4.00 × 85.0 = 340 g

(ii) If actual yield = 80 %, then mass of sodium nitrate required to obtain 64.0 g of oxygen = \( \frac{100}{80} \times 340 \) g = 425 g

[340 g of sodium nitrate produces 80 % of the desired yield

∴ 340/80 g produces 80/80 = 1 % of the desired yield

So 100 × 340/80 g produces 100 × 1 % = 100% of the desired yield

i.e. 425 g of sodium nitrate are required.]
10. (i) Equation: $C + O_2 \rightarrow CO_2$
Mole ratios: $1 \text{ mol} \quad 1 \text{ mol} \quad 1 \text{ mol}$

[Note that this is a limiting reagent type of problem where both reactants have a specified mass and one of them is in excess. The first task is to find out which is the limiting reagent and this will determine how much carbon dioxide can form.]

Moles of C = mass / molar mass of C = 1.20 / 12.01 = 0.100 mol
Moles of O$_2$ = mass / molar mass of O$_2$ = 8.00 / 32.00 = 0.250 mol

From the equation, 1 mole of C uses 1 mole of O$_2$ to produce 1 mole of CO$_2$. If all the carbon (0.100 mole) were used it would react with 0.100 mole of O$_2$, leaving (0.250 – 0.100) mole of O$_2$ in excess unreacted and form 0.100 mole of CO$_2$, so the carbon is the limiting reagent.

Mass of carbon dioxide formed = moles × molar mass of CO$_2$
$= 0.100 \times 44.0 = 4.40 \text{ g}$

(ii) From (i), the oxygen is in excess by 0.150 mole.

Mass of oxygen in excess = moles × molar mass of O$_2$ = 0.150 × 32.0 g = 4.80 g.

11. (i) Equation: $2\text{Al} + 3\text{S} \rightarrow \text{Al}_2\text{S}_3$
Mole ratios: $2 \text{ mol} \quad 3 \text{ mol} \quad 1 \text{ mol}$

[This is another limiting reagent type of problem.]

Moles of Al = mass / molar mass of Al = 10.0 / 26.98 = 0.371 mol
Moles of S = mass / molar mass of S = 15.0 / 32.07 = 0.468 mol

From the equation, 1 mole of S requires 2/3 mole of Al and produces 1/3 mole of Al$_2$S$_3$.
∴ if all the sulfur reacted, it would require $(2/3) \times 0.468$ mole of Al
$= 0.312$ mole and produce $(1/3) \times 0.468$ mole of Al$_2$S$_3$.

As 0.371 mole of Al are available, then the sulfur is the limiting reagent and the aluminium is in excess by $(0.371 – 0.312)$ mole.

Mass of Al$_2$S$_3$ produced = moles × molar mass of Al$_2$S$_3$
$= (1/3) \times 0.468 \times 150.2 = 23.4 \text{ g}$.

(ii) From (i), the aluminium is in excess by $(0.371 – 0.312) = 0.059$ mol

Mass of aluminium in excess = moles × molar mass of Al
$= 0.059 \times 26.98 \text{ g} = 1.59 \text{ g}$.
12. (i) From the given equation,

\[
2 \text{NaHSO}_3 + \text{H}_2\text{SO}_4 \rightarrow 2\text{SO}_2 + 2\text{H}_2\text{O} + \text{Na}_2\text{SO}_4
\]

Mole ratios: 2 mol : 2 mol
or 1 mol : 1 mol

In terms of mass, 1 mole = 1 molar mass or 1 gram formula weight of sodium hydrogensulfite and produces 1 mole = 1 molar mass or 1 gram formula weight of sulfur dioxide.

\[\text{mass ratios: } 104.1 \text{ g : } 64.07 \text{ g}\]

If one used tonne formula weights (wherein a mole would be defined as the amount of element or compound present in its formula weight expressed as tonnes) instead of gram formula weights, 1 tonne formula weight of sodium hydrogensulfite produces 1 tonne formula weight of sulfur dioxide.

Mass ratios in tonnes

\[104.1 \text{ tonne : } 64.07 \text{ tonne}\]

\[\therefore 1.00 \text{ tonne sodium hydrogensulfite produces } \frac{64.07}{104.1} \text{ tonne of sulfur dioxide} = 0.615 \text{ tonne.}\]

(ii) From (i), the ratio of the mass of sodium hydrogensulfite reacting to the mass of sulfur dioxide produced = 104.1 : 64.07. This ratio is independent of the mass units used as it is based entirely on the sum of the relative masses of all the atoms that constitute the two compounds - one formula unit of NaHSO_3 is 104.1/64.07 times heavier than one formula unit of SO_2 regardless of the mass unit employed. If as usual the gram was taken as the mass unit, then 104.1 g of sodium hydrogensulfite produces 64.07 g of sulfur dioxide. If tonnes were used as the mass unit, then 104.1 tonnes of sodium hydrogensulfite produces 64.07 tonnes of sulfur dioxide. If Martian globs were used as the mass unit, then 104.1 globs of sodium hydrogensulfite produces 64.07 globs of sulfur dioxide.

Thus 5.0 globs of sodium hydrogensulfite produces \((\frac{64.07}{104.1}) \times 5.0\) globs of sulfur dioxide = 3.1 globs.

13. Equation: \(\text{HNO}_3 + \text{NaOH} \rightarrow \text{NaNO}_3 + \text{H}_2\text{O}\)

Mole ratios: 1 mol : 1 mol

The equation factor = 1

\[\therefore 0.50 \text{ mole of nitric acid reacts with } 0.50 \text{ mole of sodium hydroxide.}\]

Mass of sodium hydroxide in 0.50 mole = \(0.50 \times \text{molar mass of NaOH}\)

\[= 0.50 \times 40.0 \text{ g} = 20 \text{ g.}\]
14. (a) Molar mass of Ni(NO$_3$)$_2$.6H$_2$O = 290.8 g mol$^{-1}$.
From its formula, each mole of Ni(NO$_3$)$_2$.6H$_2$O contains one mole of Ni$^{2+}$ ions.
\[ \text{moles of Ni}^{2+} = \text{moles of Ni(NO$_3$)$_2$.6H$_2$O} = \frac{\text{mass}}{\text{molar mass}} \]
\[ = \frac{29.1}{290.8} \text{ mol} = 0.100 \text{ mol} \]

(b) From the formula for nickel(II) nitrate-6 water,
moles of NO$_3^-$ = 2 × moles of Ni(NO$_3$)$_2$.6H$_2$O.
Moles of Ni(NO$_3$)$_2$.6H$_2$O = 0.100 mol
\[ \therefore \text{moles of NO}_3^- = 2 \times 0.100 = 0.200 \text{ mol} \]

(c) Number of Ni$^{2+}$ ions = $N_A \times \text{moles of Ni}^{2+} = 6.022 \times 10^{23} \times 0.100$
\[ = 6.02 \times 10^{22} \text{ Ni}^{2+} \text{ ions.} \]

(d) Number of NO$_3^-$ ions = $N_A \times \text{moles of NO}_3^- = 6.022 \times 10^{23} \times 0.200$
\[ = 12.0 \times 10^{22} \text{ NO}_3^- \text{ ions.} \]

15. (a) Equation: Zn + 2HCl → ZnCl$_2$ + H$_2$
Mole ratios: 1 mol 2 mol 1 mol 1 mol
Moles of H$_2$ produced = moles of Zn reacting.
Moles of Zn in 15.8 g = mass / molar mass of Zn = 15.8 / 65.39 = 0.242 mol
\[ \therefore \text{moles of H}_2 = 0.242 \text{ mol} \]
Mass of H$_2$ = moles of H$_2$ × molar mass of H$_2$
\[ = 0.242 \times 2.016 \text{ g} = 0.488 \text{ g}. \]

(b) Moles of ZnCl$_2$ formed = moles of Zn reacting = 0.242 mol
\[ \therefore \text{mass of ZnCl}_2 = \text{moles} \times \text{molar mass of ZnCl}_2 = 0.242 \times 136.3 \text{ g} = 33.0 \text{ g}. \]