

W1 WORKSHOP ON STOICHIOMETRY

INTRODUCTION

These notes and exercises are designed to introduce you to the basic concepts required to understand a chemical formula or equation. Relative atomic masses of all the elements can be found in the Periodic Table on the back cover of this Laboratory Handbook.

Atomic, molecular and formula weights

The smallest bit of matter that can be weighed reliably contains an enormous number of atoms and hence a convenient unit for describing a large number of atoms is required. This unit is called the **mole** and is the amount of substance that contains as many elementary units as there are carbon atoms in exactly 12 g of carbon-12. This number, approximately 6.022×10^{23} is called the **Avogadro constant** and given the symbol N_A (or less commonly, L).

One mole (1 mol) of hydrogen atoms contains 6.022×10^{23} H atoms

One mole (1 mol) of helium atoms contains 6.022×10^{23} He atoms

One mole (1 mol) of water molecules contains 6.022×10^{23} H₂O molecules

One mole (1 mol) of sodium ions contains 6.022×10^{23} Na⁺ ions

One mole (1 mol) of \$1 coins is worth $\$6.022 \times 10^{23}$

The **relative atomic mass** of any element is the mass (in gram) that contains 1 mol of atoms of that element. The relative atomic mass is often called the **atomic mass** or **atomic weight**.

eg 1. What mass of copper contains 6.022×10^{23} Cu atoms?

The relative atomic mass of Cu = 63.55

By definition, 63.55 g of copper contains 1 mol of Cu atoms;

or 63.55 g of copper contains 6.022×10^{23} Cu atoms.

eg 2. How many neon atoms are present in 10.1 g of neon?

The relative atomic mass of Ne = 20.2.

Therefore 20.2 g of neon contains 6.022×10^{23} Ne atoms.

By proportion, 10.1 g of neon contains $\frac{10.1}{20.2} \times 6.022 \times 10^{23} = 3.01 \times 10^{23}$ Ne atoms.

eg 3. The mass of one atom of an element X is found to be 2.00×10^{-23} g. What is the relative atomic mass of element X?

1 atom of X has mass 2.00×10^{-23} g

Therefore 1 mol of atoms of X has mass $(6.022 \times 10^{23}) \times (2.00 \times 10^{-23})$ g = 12.0 g.

Therefore the relative atomic mass of X = 12.0.

W1-2

Most elements exist naturally as a mixture of isotopes. Each isotope has its own characteristic atomic weight that is close to, but not exactly equal to its mass number. For example,

| isotope | % natural abundance | atomic weight |
|------------------|---------------------|----------------------------------|
| ^{12}C | 98.89 | 12.0000 (standard by definition) |
| ^{13}C | 1.11 | 13.0034 |
| ^{14}N | 99.63 | 14.0031 |
| ^{15}N | 0.37 | 15.0001 |
| ^{79}Br | 50.69 | 78.9183 |
| ^{81}Br | 49.31 | 80.9163 |

The relative atomic mass of an element is a weighted average of the relative masses of all isotopes of that element.

eg 4. Bromine contains 50.69% of the isotope ^{79}Br and 49.31% of the isotope ^{81}Br . What is the relative atomic mass of bromine?

$$\text{Relative atomic mass} = 50.69 \times \frac{78.918}{100} + 49.31 \times \frac{80.916}{100} = 79.90$$

The **relative molecular weight** of a molecule is the sum of the atomic weights of the constituent atoms multiplied by the number of those atoms present. The relative molecular weight is often called the **molecular weight**.

$$\text{H}_2 \text{ has a molecular mass} = 2 \times 1.008 = 2.016$$

$$\text{O}_2 \text{ has a molecular mass} = 2 \times 16.00 = 32.00$$

$$\text{H}_2\text{O has a molecular mass} = (2 \times 1.008) + 16.00 = 18.02$$

$$\text{CO}_2 \text{ has a molecular mass} = 12.01 + (2 \times 16.00) = 44.01$$

The **formula weight** of an ionic compound is the sum of the atomic masses of the constituent atoms multiplied by the number of those atoms present in the empirical formula.

$$\text{NaCl has a formula weight} = 22.99 + 35.45 = 58.44$$

$$\text{MgCl}_2 \text{ has a formula weight} = 24.31 + (2 \times 35.45) = 95.21$$

The general expression relating the number of moles to mass is:

$$\text{Amount of substance (in mol)} = \frac{\text{mass of substance}}{\text{formula weight}}$$

or

$$n = \frac{m}{M}$$

eg 5. What is the mass of 1.50 mol of carbon dioxide?

$$\text{The molecular weight of CO}_2 = 12.01 + (2 \times 16.00) = 44.01.$$

$$\text{Therefore 1.00 mol of CO}_2 \text{ has mass of } 44.01 \text{ g}$$

$$\text{Therefore 1.50 mol of CO}_2 \text{ has mass of } 44.01 \times 1.50 = 66.0 \text{ g}$$

eg 6. What is the mass of 0.270 mol of the salt lithium fluoride?

Lithium fluoride has the formula LiF.

$$\text{Therefore 1.00 mole of LiF has mass } (6.941 + 19.00) \text{ g} = 25.94 \text{ g.}$$

$$\text{Therefore 0.270 mole of LiF has mass } 0.270 \times 25.94 \text{ g} = 7.00 \text{ g.}$$

eg 7. What amount of calcium bromide is present in 6.76 g of CaBr_2 ?

The formula weight of $\text{CaBr}_2 = 40.08 + (2 \times 79.90) = 199.88$

Therefore 199.88 g of $\text{CaBr}_2 = 1.00$ mol

Therefore 6.76 g of $\text{CaBr}_2 = 1.00 \times \frac{6.76}{199.88}$ mol = 0.0338 mol

eg 8. How many molecules of H_2O are present in 0.500 mol of water?

1 mol of water contains 6.022×10^{23} H_2O molecules

Therefore 0.500 mol contains $0.500 \times (6.022 \times 10^{23}) = 3.01 \times 10^{23}$ H_2O molecules.

eg 9. How many molecules are present in 2.00 g of table sugar, $\text{C}_{12}\text{H}_{22}\text{O}_{11}$?

The relative molecular mass of $\text{C}_{12}\text{H}_{22}\text{O}_{11} = (12 \times 12.01) + (22 \times 1.008) + (11 \times 16.00)$
 $= 342.30$

Therefore 2.00 g = $\frac{2.00}{342.30}$ mol

Therefore 2.00 g contains $\frac{2.00}{342.30} \times 6.022 \times 10^{23}$ molecules of $\text{C}_{12}\text{H}_{22}\text{O}_{11}$
 $= 3.52 \times 10^{21}$ molecules.

Gases

The gaseous elements hydrogen, nitrogen, oxygen, fluorine and chlorine occur as diatomic molecules. That is, they exist naturally as H_2 , N_2 , O_2 , F_2 and Cl_2 rather than as the atomic species. (Bromine, Br_2 , and iodine, I_2 , though not gases are also diatomic elements.) In contrast, the noble gases helium, neon, argon, krypton, xenon and radon all exist as monatomic species.

1 mole of any gas occupies 22.4 L at standard temperature and pressure.

Standard temperature and pressure (STP) is defined as 0 °C and 1 atmosphere.

1.00 mol of hydrogen (H_2 molecules) occupies 22.4 L at STP.

1.00 mol of helium (He atoms) occupies 22.4 L at STP.

22.4 L of chlorine (Cl_2 molecules) contains 1.00 mol of Cl_2 molecules at STP.

11.2 L of fluorine (F_2 molecules) contains 0.500 mol of F_2 molecules at STP.

22.4 L of neon contains 6.02×10^{23} atoms of Ne at STP.

eg 10. What is the volume of 76.0 g of fluorine at STP?

The molecular weight of $\text{F}_2 = 2 \times 19.00 = 38.00$

Amount of $\text{F}_2 = \frac{76.00}{38.00} = 2.00$ mol

Therefore volume = 2.00 mol $\times 22.4$ L mol⁻¹ = 44.8 L.

eg 11. What is the mass of 105 L of oxygen at STP?

The molecular weight of $\text{O}_2 = 2 \times 16.00 = 32.00$

105 L of $\text{O}_2 = \frac{105}{22.4}$ mol of $\text{O}_2 = \frac{105}{22.4}$ mol $\times 32$ g mol⁻¹ = 150 g.

W1-4

Q1. Calculate the mass of 2.0 mol of silicon.

Q2. Calculate the mass of 0.37 mol of barium chloride.

Q3. Calculate the amount of S atoms (in mol) present in 2.8 g sulfur.

Q4. Calculate the amount of H₂O molecules (in mol) present in 36.0 g of water.

Q5. Calculate the mass of 6.022×10^{23} molecules of hydrogen.

Q6. Calculate the amount of CO_2 molecules (in mol) present in 2.0×10^{20} molecules of carbon dioxide.

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Q7. Calculate the amount of Ar atoms (in mol) present in 5.6 L of argon at STP.

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Q8. Calculate the mass of 50.0 L of nitrogen gas at STP.

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Q9. Calculate the atomic weight and the molecular weight of a natural sample of chlorine, which contains the isotopes: ^{35}Cl (at. wt. 34.97, 75.77%) and ^{37}Cl (at. wt. 36.97, 24.23%).

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Structural, molecular and empirical formulas

The **structural formula** indicates the arrangement of the atoms and the bonds within the molecule, *eg* H–O–O–H for hydrogen peroxide.

The **molecular formula** shows the actual number of atoms of each element in a molecule of the compound, *eg* H₂O₂ for hydrogen peroxide.

The **empirical formula** of a compound is the simplest integer ratio of the elements in that compound, *eg* HO for hydrogen peroxide.

Metals, ionic compounds, some non-metallic elements and many inorganic compounds consist of “infinite” lattices of regularly repeating units. The term “molecular weight” has no meaning when applied to such compounds, there being no molecules. For these compounds, the term **formula weight** should be used, where the formula is the empirical formula. (See page W1-2.)

Percent composition of a compound

Each pure compound always has the same composition by weight, regardless of how it is produced. For example, glucose, C₆H₁₂O₆, always has 6 C atoms, 12 H atoms and 6 O atoms in every one of its molecules. Given the atomic weights of the component elements, the % by weight for each element in the compound can be calculated.

eg 12. Calculate the % composition of glucose.

The molar mass (gram molecular weight) of glucose, C₆H₁₂O₆,
 $= (6 \times 12.01 + 12 \times 1.01 + 6 \times 16.00) = 180.18 \text{ g mol}^{-1}$

From the formula of glucose, 1 molecule of glucose contains 6 atoms of C.

Hence 1 mol of glucose would contain 6 mol of C.

As 1 mole of C weighs 12.01 g,

then the mass of C in 1 mol of glucose = $6 \times 12.01 \text{ g} = 72.06 \text{ g}$.

$$\% \text{ C in C}_6\text{H}_{12}\text{O}_6 = \frac{72.06}{180.18} \times 100\% = 39.99\% \text{ by mass.}$$

Similarly, the percentage of H and O can be calculated as follows:

Each mole of glucose contains 12 mol of H and 6 mol of O, respectively.

$$\% \text{ H} = \frac{12 \times 1.008}{180.18} \times 100\% = 6.71\% \text{ H}$$

$$\% \text{ O} = \frac{6 \times 16.00}{180.18} \times 100\% = 53.28\% \text{ O.}$$

[Note that the sum of the % of all the elements must add up to 100%.]

Recall that the empirical formula of a compound is the simplest integer ratio of the elements in that compound. Given the % composition of a compound, its empirical formula can be calculated. The processes involved are simply the reverse of those just done above.

eg 13. Determine the empirical formula of a compound that returns the following data for % by mass: iron (Fe): 63.5 %, sulfur (S): 36.5 %

From the data, in 100 g of compound, there would be 63.5 g of Fe combined with 36.5 g of S. To obtain the empirical formula, we need to determine the relative amounts of Fe and S present. Because Fe and S have different atomic weights, to calculate the number of moles of Fe and of S atoms present in the compound, it is necessary to divide the mass of each by the atomic weight of the element.

$$\begin{aligned}\text{Amount of Fe} &= \frac{63.5}{55.85} = 1.137 \text{ mol of Fe} \\ \text{Amount of S} &= \frac{36.5}{32.07} = 1.138 \text{ mol of S}\end{aligned}$$

Thus the ratio of amount of Fe to amount of S in the compound is

$$1.137 \text{ mol of Fe} : 1.138 \text{ mol of S}$$

This is also the ratio of the number of atoms of Fe to the number of atoms of S. However, the empirical formula must have integer quantities for all the numbers of atoms in it. In this example, it is obvious that, within the usual allowable experimental error in analytical data of 0.3 %, the ratio of atoms of Fe: atoms of S = 1.00 : 1.00.

Thus the empirical formula is FeS.

eg 14. Determine the empirical formula of an unknown compound that has the following % composition by mass: nitrogen 26.2%, chlorine 66.4%, hydrogen 7.50%

[Note that these percentages total to 100%, within the experimental error. If they do not total to 100%, the difference is attributed to oxygen, whose analysis is difficult.]

$$\begin{aligned}\text{N} : \text{Cl} : \text{H} &= \frac{\% \text{ N}}{\text{atomic weight N}} : \frac{\% \text{ Cl}}{\text{atomic weight Cl}} : \frac{\% \text{ H}}{\text{atomic weight H}} \\ &= \frac{26.2}{14.01} : \frac{66.4}{35.45} : \frac{7.50}{1.008} * \\ &= 1.870 : 1.873 : 7.440 \\ &= 1.00 : 1.00 : 3.98 \\ &\approx 1 : 1 : 4\end{aligned}$$

In order to convert these numbers to integers they must be divided by the highest common factor. A good starting point is to divide each by the smallest number, in this case 1.870.

Thus the empirical formula is NClH₄.

*Note that you must divide each mass by the atomic weight of the element.

[A common error is to divide by the molecular weight for species that occurs in nature as diatomic molecules, eg 28.02 for N₂, 70.90 for Cl₂, 32.00 for O₂ or 2.016 for H₂.]

W1-8

Q10. Determine the percentage by weight of bromide ion in potassium bromide (KBr).

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Q11. An iron ore has the composition of 70.0% Fe and 30.0% O by mass. What is the empirical formula of the ore?

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Q12. An organic compound containing only carbon, hydrogen and oxygen returns the % mass analysis: C 64.9 %; H 13.5 %. What is its empirical formula?

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Chemical Equations

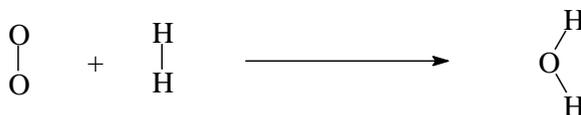
A chemical equation describes the changes occurring in a reaction. For example, the reaction between hydrogen gas and oxygen gas produces water. In words, the reaction may be written as:



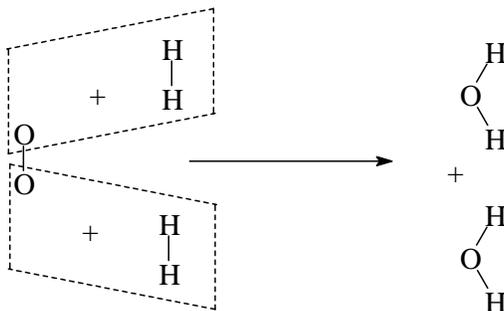
The standard symbols are used to represent the chemical species involved. As explained on page W1-3, hydrogen and oxygen are diatomic gases - they have the molecular formulas H_2 and O_2 , respectively. Water is composed of two hydrogen atoms and one oxygen atom, *i.e.* it has the molecular formula H_2O . The reaction may be represented as:



But does 1 hydrogen molecule react with 1 oxygen molecule to produce 1 water molecule? A quick look at the following diagram shows that the answer is “no” - one oxygen atom has just ‘disappeared’.



Atoms are neither created nor destroyed in chemical reactions (The Law of Conservation of Mass). Consequently there must be the same number of atoms of any particular type on both sides of the equation, irrespective of how they are arranged. This can be achieved in our current example if two molecules of hydrogen react with one molecule of oxygen to produce two molecules of water.



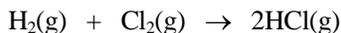
or in symbols



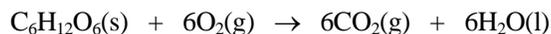
This process of ensuring that the same number of atoms of any particular type appear on both sides of the equation is called **balancing the equation**. A number used as a prefix indicates the ratio of reacting species in the equation. In this case *two* hydrogen molecules react with *one* oxygen molecule to form *two* water molecules. As the numerical prefixes represent ratios, it is just as valid (and much more usual) to say that *two mole* of hydrogen reacts with *one mole* of oxygen to form *two mole* of water. The physical *state* of the various species is indicated by (s) for a solid, (l) for a liquid, (g) for a gas or (aq) for an aquated species. In this example the water is generated as a gas and the final balanced equation is written:



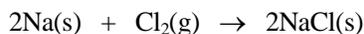
eg 15. Hydrogen and chlorine react to form the compound hydrogen chloride, all species being gases. Write the balanced molecular equation.



eg 16. The carbohydrate, glucose, is oxidised to carbon dioxide and water in biological systems. Write the overall molecular reaction.



eg 17. Solid sodium metal reacts with chlorine gas to form sodium chloride. Give the equation.



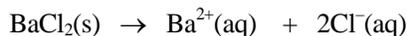
In this example, note how the empirical formulas, Na for the metal sodium and NaCl for the ionic compound sodium chloride, are used.

Reactions in solution

When an ionic solid dissolves in water to form a solution, the charges on the ions are indicated.

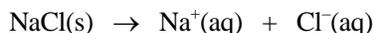
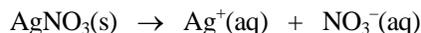


Again, the atoms must balance. Notice also that the electrical charges present on both sides of the equation must balance as well. In another example:

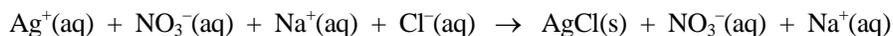


In words, this means that 1 mole of barium chloride dissolves in water to produce 1 mole of barium ions and 2 mole of chloride ions.

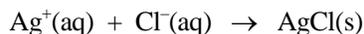
Silver nitrate and sodium chloride are both soluble compounds, and in water solution release their component ions, shown in the following ionic equations.



Silver chloride is insoluble in water. When a solution of silver nitrate is mixed with a solution of sodium chloride, the $\text{Ag}^+(\text{aq})$ ions react with the $\text{Cl}^-(\text{aq})$ ions to *precipitate* AgCl as shown in the following ionic equation.



Note that only the Ag^+ and the Cl^- ions have reacted, leaving the Na^+ and NO_3^- ions free in the solution. As these last two ions have not in fact entered into a reaction, they should be deleted from the equation in much the same way as common terms are cancelled from both sides of a mathematical equation. Such ions are called **spectator ions**. [Initially it may be helpful to write down all the ions that are being mixed together in order to establish whether any combination can form an insoluble salt, and then cancel out the spectator species.] The net ionic equation is thus:



Q13. Balance each of the following molecular equations:

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| Carbon is burnt in a limited supply of oxygen to produce carbon monoxide. $\text{C(s)} + \text{O}_2\text{(g)} \rightarrow \text{CO(g)}$ |
| Nitrogen and hydrogen react to produce ammonia. $\text{N}_2\text{(g)} + \text{H}_2\text{(g)} \rightarrow \text{NH}_3\text{(g)}$ |
| Sodium metal reacts with bromine to form sodium bromide. $\text{Na(s)} + \text{Br}_2\text{(l)} \rightarrow \text{NaBr(s)}$ |
| Iron reacts with oxygen to form iron(III) oxide. $\text{Fe(s)} + \text{O}_2\text{(g)} \rightarrow \text{Fe}_2\text{O}_3\text{(s)}$ |

Q14. Complete the following table. (See page E2-2 if you need help.)

| Formula | Name | Formula | Name |
|-----------------------------|-----------------|------------------------------|-----------------------|
| OH^- | | | acetate ion |
| | nitrite ion | CN^- | |
| | nitrate ion | HS^- | |
| $\text{C}_2\text{O}_4^{2-}$ | | | permanganate ion |
| | perchlorate ion | | hydrogencarbonate ion |
| | carbonate ion | | phosphate ion |
| $\text{S}_2\text{O}_3^{2-}$ | | H_2PO_4^- | |
| | sulfate ion | | ammonium ion |
| | sulfite ion | $\text{Cr}_2\text{O}_7^{2-}$ | |

Q15. Indicate the charges on the ions and balance the following ionic equations:

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| Potassium iodide is dissolved in water. $\text{KI(s)} \rightarrow \text{K (aq)} + \text{I (aq)}$ |
| Sodium carbonate is dissolved in water. $\text{Na}_2\text{CO}_3\text{(s)} \rightarrow \text{Na (aq)} + \text{CO}_3\text{ (aq)}$ |
| Ammonium chloride is dissolved in water. $\text{NH}_4\text{Cl(s)} \rightarrow \text{NH}_4\text{ (aq)} + \text{Cl (aq)}$ |
| Calcium hydroxide is dissolved in water. $\text{Ca(OH)}_2\text{(s)} \rightarrow \text{Ca (aq)} + \text{OH (aq)}$ |

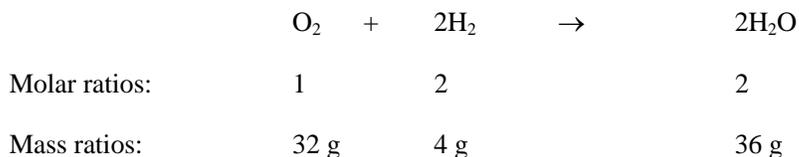
Q16. Write the ionic equations for the reactions that occur when solid sodium carbonate and solid calcium chloride dissolve in water. Also write the ionic equation for the precipitation of calcium carbonate resulting from mixing the two solutions.

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| Dissolution of sodium carbonate |
| Dissolution of calcium chloride |
| Mixing above solutions |

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| Demonstrator's Initials | |
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Limiting Reactant

A balanced chemical equation indicates the relative amounts of the chemicals that take part in a reaction. It bears absolutely no relationship whatsoever to the amounts of chemicals that a chemist might decide to mix together in a vessel. For instance



If 32 g of O_2 is mixed with 4 g of H_2 and the mixture ignited, 36 g of H_2O will form. But what is the result if 16 g of O_2 and 4 g of H_2 are added to the reaction vessel? The oxygen will react completely to give 18 g of water and there will still be 2 g of H_2 left. In this case, oxygen is said to be the limiting reactant and hydrogen is said to be in excess.

eg 18 Ammonia is oxidised at high temperature in the first step of the Ostwald process to produce nitric acid. The products of the reaction are nitric oxide and water vapour. If a 51.10 g sample of NH_3 is reacted with 64.00 g of O_2 , what mass of the various products and reactants are present after the reaction is completed?

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| Balanced equation | $5\text{O}_2(\text{g})$ | + | $4\text{NH}_3(\text{g})$ | \rightarrow | $4\text{NO}(\text{g})$ | + | $6\text{H}_2\text{O}(\text{g})$ |
| Molar ratios | 5 | | 4 | | 4 | | 6 |
| Initial amounts (in mol) | $\frac{64.00}{32.00} = 2.00$ | | $\frac{51.10}{17.03} = 3.00$ | | 0 | | 0 |
| change* | -2.00 | | -1.60 | | +1.60 | | +2.40 |
| final amounts (in mol) | 0 | | 1.40 | | 1.60 | | 2.40 |
| final amounts (in g) | 0 | | 1.40×17.03 | | 1.60×30.01 | | 2.40×18.02 |
| = | 0 | | 23.84 | | 48.02 | | 43.25 |

* Mole ratio of $\text{NH}_3 : \text{O}_2$ from equation = $4 : 5 = 0.8$

Mole ratio of $\text{NH}_3 : \text{O}_2$ from above calculations $3 : 2 = 1.5$

That is, excess NH_3 is present and thus O_2 is the limiting reagent.

(If the $\text{NH}_3 : \text{O}_2$ ratio had been less than 0.8, O_2 would have been present in excess and NH_3 would have been the limiting reagent.)

Having established that O_2 is the limiting reactant and will all be consumed, the change in reactants and products is calculated from the equation as follows:

5 mol of O_2 reacts with 4 mol of NH_3 to give 4 mol of NO and 6 mol of H_2O .

\therefore 2 mol of O_2 reacts with $\frac{4 \times 2}{5}$ mol of NH_3 to give $\frac{4 \times 2}{5}$ mol of NO and $\frac{6 \times 2}{5}$ mol of H_2O .

\therefore 2×32.00 g of O_2 reacts with $\frac{4 \times 2}{5} \times 17.03$ g of NH_3 to give

$\frac{4 \times 2}{5} \times 30.01$ g of NO and $\frac{6 \times 2}{5} \times 18.02$ g of H_2O .

i.e. 64.00 g of O_2 reacts with 27.25 g of NH_3 to give 48.02 g of NO and 43.25 g of H_2O . No O_2 is left over, but 23.84 g of the NH_3 remains unreacted.

Solutions

The concentration of a solute in a solution may be expressed as a *molarity*. Concentration is indicated by square brackets around the species to which the concentration refers and has units of mol L^{-1} or M.

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| $\text{Concentration (in mol L}^{-1}\text{)} = \frac{\text{amount (in mol)}}{\text{volume (in L)}}$ |
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eg 19. What is the concentration of one litre of glucose solution containing 2 moles of glucose?

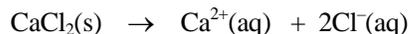
$$[\text{Glucose}] = \frac{2 \text{ mol}}{1 \text{ L}} = 2 \text{ mol L}^{-1} = 2 \text{ M}$$

eg 20. One mole of sodium hydroxide is dissolved in water and made up to 250 mL. What is the concentration of the solution?

$$250 \text{ mL} = 250 \times 10^{-3} \text{ L} = 0.250 \text{ L}$$

$$\text{concentration of NaOH in 1 L} = \frac{1 \text{ mol}}{0.250 \text{ L}} = 4 \text{ mol L}^{-1} = 4 \text{ M}$$

eg 21. A solution of calcium chloride has a concentration of 0.30 M. What is the concentration of the individual ions in the solution?



That is, 1 mol of $\text{CaCl}_2(\text{s})$ produces 1 mol of $\text{Ca}^{2+}(\text{aq})$ and 2 mol of $\text{Cl}^{-}(\text{aq})$ ions.

Thus, if concentration of $\text{CaCl}_2 = 0.30 \text{ M}$, from the equation it can be seen that

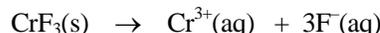
$$[\text{Ca}^{2+}] = 0.30 \text{ M} \text{ and } [\text{Cl}^{-}] = 2 \times 0.30 \text{ M} = 0.60 \text{ M}.$$

eg 22. Chromium(III) fluoride (6.66 g) is dissolved in water and the volume made up to 400.0 mL. What is the concentration of the ions in the solution.

$$\text{Relative formula weight of CrF}_3 = 54.00 + (3 \times 19.00) = 111.00$$

$$\text{Amount of CrF}_3 = \frac{6.66}{111.00} \text{ mol} = 0.0600 \text{ mol}$$

$$\text{concentration of CrF}_3 = \frac{0.0600 \text{ mol}}{0.400 \text{ L}} = 0.150 \text{ mol L}^{-1} = 0.150 \text{ M}$$



Thus, if concentration of $\text{CrF}_3 = 0.150 \text{ M}$, from the equation it can be seen that

$$[\text{Cr}^{3+}] = 0.150 \text{ M} \text{ and } [\text{F}^{-}] = 3 \times 0.150 \text{ M} = 0.450 \text{ M}.$$

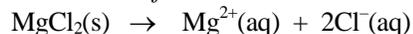
eg 23. What volume of solution is required to form a 0.040 M solution of sodium chloride from 0.0010 moles of the salt?

$$\text{Concentration (mol L}^{-1}\text{)} = \frac{\text{amount (in mol)}}{\text{volume (in L)}}$$

$$\text{Therefore volume (in L)} = \frac{\text{amount (in mol)}}{\text{concentration (in M)}}$$

$$\text{Therefore volume (in L)} = \frac{0.0010 \text{ mol}}{0.040 \text{ M}} = 0.025 \text{ L} = 25 \text{ mL}$$

eg 24. What amount of chloride ion is in 150 mL of a 0.50 M solution of magnesium chloride?

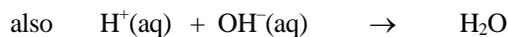
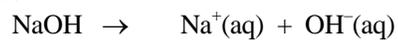


$$[\text{Mg}^{2+}(\text{aq})] = 0.50 \text{ M} \text{ and } [\text{Cl}^{-}(\text{aq})] = 2 \times 0.50 \text{ M} = 1.00 \text{ M}$$

$$\text{Concentration (in mol L}^{-1}\text{)} = \frac{\text{amount (in mol)}}{\text{volume (in L)}}$$

$$\begin{aligned} \text{Therefore amount (in mol)} &= \text{Concentration (in mol L}^{-1}\text{)} \times \text{volume (in L)} \\ &= 1.00 \text{ mol L}^{-1} \times 0.150 \text{ L} = 0.150 \text{ mol} \end{aligned}$$

eg 25. What volume of 0.11 M hydrochloric acid will react with 25 mL of 0.090 M sodium hydroxide solution?



$$[\text{OH}^-] = 0.090 \text{ M} \quad \text{and } 25 \text{ mL} = 0.025 \text{ L}$$

$$\begin{aligned} \text{moles of OH}^- = \text{moles of H}^+ &= 0.090 \text{ mol L}^{-1} \times 0.025 \text{ L} \\ &= 0.00225 \text{ mol} \end{aligned}$$

$$\begin{aligned} \therefore \text{Volume of H}^+ \text{ needed} &= \frac{0.00225 \text{ mol}}{0.11 \text{ mol L}^{-1}} \\ &= 0.02045 \text{ L} \\ &= 20. \text{ mL} \end{aligned}$$

Q17. Calculate the mass of sodium carbonate ($\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$) required to make 250. mL of a 0.100 M solution.

Q18. What mass of barium sulfate will be precipitated when 125 mL of a 0.20 M solution of barium chloride is mixed with 200. mL of a 0.17 M solution of sodium sulfate. (Hint: work out which reagent is limiting.)

W1-16

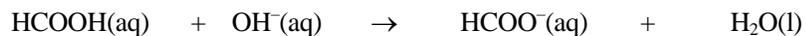
Q19. Pure formic acid (HCOOH), is a liquid monoprotic acid decomposed by heat to carbon dioxide and hydrogen, according to the following equation:



(i) The density of formic acid is 1.220 g mL^{-1} . How many moles of HCOOH are in 1.000 L of pure formic acid?

(ii) What mass of pure formic acid should be diluted to 1.00 L to form a 2.00 M solution?

(iii) What volume of 0.250 M sodium hydroxide solution would react with 30.0 mL of this dilute solution of formic acid, according to the following equation?



(iv) What is the maximum volume of carbon dioxide at STP that could be obtained by heating 1.0 mol of formic acid?

(v) How many molecules of carbon dioxide would it contain?

Q20. Consider the reaction $4\text{Al(s)} + 3\text{O}_2\text{(g)} \rightarrow 2\text{Al}_2\text{O}_3\text{(s)}$

Identify the limiting reagent in each of the following reaction mixtures. What mass of $\text{Al}_2\text{O}_3\text{(s)}$ will be produced in each case?

1.0 mol Al and 1.0 mol O_2

0.75 mol Al and 0.50 mol O_2

75.89 g Al and 112.25 g O_2

51.28 g Al and 48.22 g O_2

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| Demonstrator's Initials | |
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POST-WORK

The following questions are further examples of the basic types of stoichiometric calculations. Proficiency in these types of calculations is essential for any serious student of Chemistry. If time permits they may be done in your laboratory session, otherwise they should be done before your next laboratory session. Worked answers to all these problems are available from WebCT.

Q1. Write the equation that relates amount of a substance to mass.

Q2. Calculate the mass of 1.87 mol of sulfur trioxide.

Q3. Calculate the amount (in mol) present in 200.0 g of silicon tetrachloride.

Q4. Calculate the mass of 2.00×10^{20} molecules of water.

Q5. Calculate the volume (in L) present in 5.45×10^{22} atoms of helium at STP.

Q6. Calculate the relative atomic mass of a natural sample of zinc, which contains the isotopes with masses and abundances given:

| isotope | atomic weight | abundance | isotope | atomic weight | abundance |
|------------------|---------------|-----------|------------------|---------------|-----------|
| ^{64}Zn | 63.929 | 48.6% | ^{68}Zn | 67.925 | 18.8% |
| ^{66}Zn | 65.926 | 27.9% | ^{70}Zn | 69.925 | 0.6% |
| ^{67}Zn | 66.927 | 4.1% | | | |

Q7. An iron supplement is used to treat anaemia and 50 mg (*i.e.* 50×10^{-3} g) of Fe^{2+} is required per tablet. If the iron compound used in the tablet is $\text{FeSO}_4 \cdot 7\text{H}_2\text{O}$, what mass of this compound would be required per tablet to provide the desired amount of Fe^{2+} ?

Q8. Write the equation that relates concentration of a solution to amount of solute and volume of solution.

Q9. Write the net ionic equation for the reaction that occurs when a solution of barium nitrate is mixed with a solution of sodium sulfate. A white precipitate of barium sulfate forms.

Q10. One of the components of petrol is octane, C_8H_{18} . (i) Write the balanced equation for the complete combustion of octane to form carbon dioxide gas and liquid water.

(ii) What amount (in mol) of carbon dioxide is formed when 5.5 mol (1 L) of octane is burnt?

(iii) What volume of carbon dioxide would this represent at STP?

Q11. Hydrogen iodide gas (5.0 L at STP) is dissolved in water and the volume made up to 1.0 L. What is the molarity of the solution?

Q12. What volume of 0.200 M hydrochloric acid would be needed to react completely with a mixture of 0.500 g of sodium hydroxide and 0.800 g of potassium hydroxide?

Q13. A solution was prepared by dissolving nickel (II) nitrate-6-water, $\text{Ni}(\text{NO}_3)_2 \cdot 6\text{H}_2\text{O}$, (29.1 g) in some water and making the volume up to 1.00 L with water. Assuming complete dissociation of the solid into ions, calculate:

(i) The amount (in mol) of nickel(II) ions in 100. mL of this solution.

(ii) The amount (in mol) of nitrate ions in 100. mL of this solution.

(iii) The number of individual nickel(II) ions in 100. mL of solution.

Q14. What volume of 0.010 M silver nitrate solution will exactly react with 20. mL of 0.0080 M sodium chloride solution?