CHEMISTRY 91910 (2.1)

Internally assessed
4 credits

Carry out a practical investigation into a substance present in a consumer product using quantitative analysis

A full description of Achievement Standard 91910 (Chemistry 2.1) can be found at ESA **RESOURCES** at www.esa.co.nz

Chemistry terms

Below are some chemistry terms that you should be familiar with.

atom	the smallest particle of a chemical element (shows the chemical properties of that element) e.g. an atom of sodium (symbol Na); an atom of oxygen (symbol O)
molecule	two or more atoms (the same or different atoms) which have chemically combined to form a single unit of a substance e.g. oxygen molecule O_2, carbon dioxide molecule CO_2
ion	a charged particle, formed from one or more atoms, that has lost or gained one or more electrons e.g. sodium ion Na^+; oxide ion O^{2-}; carbonate ion CO_3^{2-}
element	a pure substance made up of a single type of atom e.g. sodium metal, Na; oxygen gas, O_2
compound	a pure substance made up of more than one type of atom e.g. carbon dioxide CO_2; hydrochloric acid HCl

Quantitative chemistry

Quantitative chemistry is finding out:

- how much is present in a chemical substance, or
- how much is involved in a chemical reaction, for one or more chemicals.

Before considering a practical investigation involving quantitative analysis, it is important to understand some of the *principles* of quantitative chemistry.

Significant figures indicate the level of accuracy of the data and/or apparatus. In calculations, final answers typically include three significant figures, and no rounding should occur until the final answer.

Even small amounts of substances contain very large numbers of particles, e.g. a cup of water contains about 8.3×10^{24} water molecules. Although the number of particles handled in chemistry is very large, the masses of each of these particles are usually very small. The mass of a chlorine molecule is about 1.18×10^{-22} g. It would take about 10^{22} chlorine molecules to equal a mass of just 1 g. Since these masses are small, the unit **gram** (symbol **g**) is used.

One cube of sugar contains about 1.4×10^{22} molecules of glucose.

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Calculations using amount of substance

The mole

When particularly large number of objects are needed, quantities are defined which suit these large numbers. In chemistry, the quantity chosen to describe the amount of a substance is the **mole**, abbreviated to **mol** when used as a unit. The amount of substance (in mol) is represented by the symbol *n* (thus $n(H_2O)$ means the amount of water, in moles).

The mole is defined as 'the amount of substance which contains the same number of particles (atoms, ions or molecules) as there are atoms in exactly 12 g of carbon-12 (^{12}C or C-12)'.

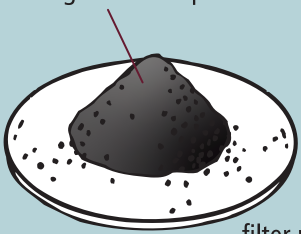
Examples

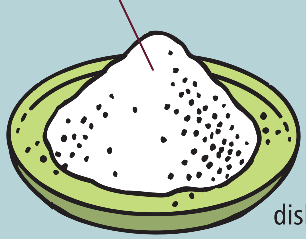
12 g of carbon is 1 mol of carbon.

The number of carbon atoms in 12 g of carbon is: 602 000 000 000 000 000 000 000, or 6.02×10^{23} atoms.

58.5 g of sodium chloride is 1 mol of sodium chloride.

The number of sodium chloride ion pairs in 58.5 g is 6.02×10^{23}.





Avogadro's number

A mole always contains 6.02×10^{23} particles. This is called **Avogadro's number, N_A**

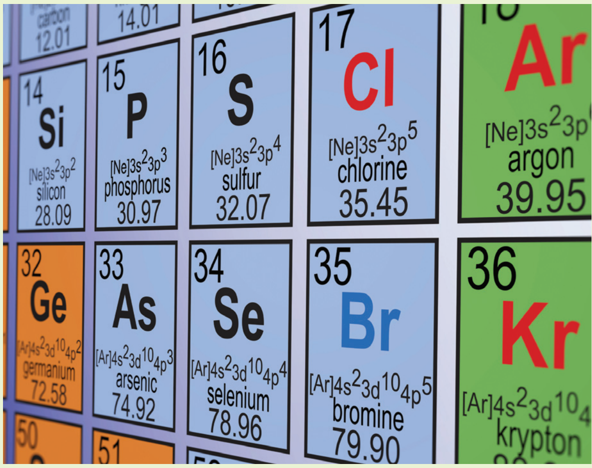
$$N_A = 6.02 \times 10^{23} \text{ mol}^{-1}$$

Avogadro's number is sometimes approximated to 6×10^{23} mol^{-1}.

Molar mass, *M*

- The mass of one mole of a substance (element, compound or ion) is the **molar mass, *M***. Molar mass has the unit **g mol^{-1}**.
- The molar mass of any substance is the sum of the molar masses of the atoms (or ions) in the formula.
- The molar mass of a simple ion (e.g. $M(Na^+)$ or $M(O^{2-})$) can be taken as the same as the molar mass of its 'parent' atom (e.g. $M(Na^+) = M(Na)$; $M(O^{2-}) = M(O)$).

Molar mass values of elements can be obtained from the periodic table at the back of this book.



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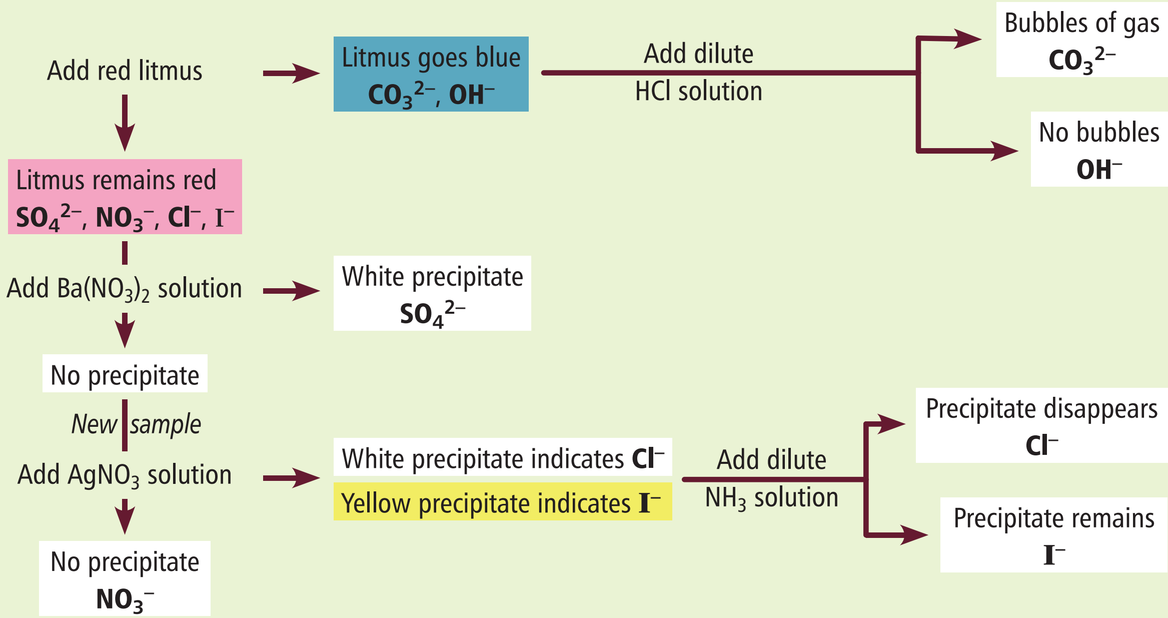
Identifying anions in aqueous solutions



Identifying ions

Answers p. 310

1. Refer to the flow chart for identifying cations in aqueous solutions (on page 62) to answer the following.

   a. i. Which cation in solution does not form a precipitate when 2 drops of dilute NaOH solution are added?

   ii. Which cations in solution form precipitates that will dissolve in excess NaOH solution?

   iii. Identify the cation in solution that forms a white precipitate when 2 drops NaOH solution are added and the precipitate then dissolves in ammonia solution.

   iv. Identify the cation in solution that forms a coloured precipitate when 2 drops NaOH solution are added and the precipitate then dissolves in NH_3 solution to produce a coloured solution.

   v. Which cations form a white precipitate when 2 drops dilute NaOH solution are added and also form a white precipitate when dilute H_2SO_4 is added to a fresh solution of the metal ion?

   b. Write ionic equations for:

   i. all cations that form a white precipitate with 2 drops of dilute NaOH solution

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ii. a cation forming a green precipitate with 2 drops of dilute NaOH solution.

2. Refer to the flow chart for identifying anions in aqueous solutions (on page 63) to answer the following.
   a. i. Identify the anion in solution that forms a white precipitate with $Ba(NO_3)_2$ solution.

      ii. Identify the anion in solution that forms a white precipitate with $AgNO_3$ solution.

      iii. Identify the anion in solution that forms a pale yellow precipitate with $AgNO_3$ solution.

      iv. Identify the anions in solution that will turn red litmus paper blue.

   b. Write an ionic equation for:
      i. an anion forming a white precipitate with $Ba(NO_3)_2$ solution

      ii. an anion forming a pale yellow precipitate with $AgNO_3$ solution.

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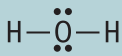
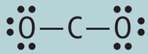
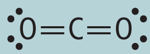
Rules for drawing Lewis structures

The total number of electrons in a Lewis structure must equal the sum of the valence electrons for all the atoms in the molecule.

The following steps show how to draw Lewis structures.

Examples

	Water, H_2O	Carbon dioxide, CO_2
Step 1 Count the total number of valence electrons present.	2 × H atoms (each with 1 valence e$^-$) 2e$^-$ 1 × O atom 6e$^-$ Total valence e$^-$ = 8	2 × O atoms (each with 6 valence e–) 12e– 1 × C atom 4e$^-$ Total valence e$^-$ = 16
Step 2 Connect the atoms with a single bond pair. Each bond represents 2 electrons.	H—O—H or H:O:H	O—C—O or O:C:O
Step 3 Place the valence electrons in pairs around the atoms, beginning with the outer atoms. (Remember: A full shell for H has 2 electrons; for other atoms, a full shell has 8 atoms.)		 Structure shows 16 electrons, but while the O atoms have 8 electrons, the C atom only has 4 electrons. To complete the 'octet' on C, move a pair of electrons from each O atom so that there are two double bonds.
Step 4 Check that the structure has the correct number of valence electrons and that the valence shells of all atoms are complete (octet rule). If they are not, move non-bonding pairs to form double or triple bonds.	The central O atom has 8 electrons – a full valence shell. The H atoms each have 2 electrons.	 Both the central C atom and the two O atoms now have 8 valence electrons.

There are some molecules that have fewer than eight electrons around the central atom. Common examples are molecules containing boron and beryllium, e.g. BH_3.

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Lewis structures

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Draw Lewis structures for:

1. the fluorine molecule

2. the silicon hydride, SiH_4, molecule

3. the sulfur dioxide, SO_2, molecule

4. the ethene, C_2H_4, molecule

5. the ethyne, C_2H_2, molecule

6. the nitrogen trichloride, NCl_3, molecule

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Summary of responses of equilibrium system to changes

Change in conditions		Direction of change in equilibrium position
Concentration	Increase products	In the reverse direction
	Decrease products	In the forward direction
	Increase reactants	In the forward direction
	Decrease reactants	In the reverse direction
Pressure	Increase	In the direction forming the least number of moles of gas
	Decrease	In the direction forming the greater number of moles of gas
Temperature	Increase	In the direction of the endothermic reaction
	Decrease	In the direction of the exothermic reaction
Catalyst added		No change in equilibrium position or in K_c; equilibrium is reached more quickly (i.e. reaction rate changes).

Determining whether a system is at equilibrium

The reaction quotient, Q, is an expression showing the ratio of reactants and products present in a system. Q uses the same expression as the equilibrium constant except that Q can have any value under any set of conditions.

If $Q = K_c$, then the reaction is at equilibrium.

If $Q > K_c$, then the reaction will proceed in the reverse direction in order to reach equilibrium.

If $Q < K_c$, then the reaction will proceed in the forward direction to reach equilibrium.

Example

The value of the K_c for the following reaction is 640 at 25 °C:

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$$

A mixture of the gases has concentrations given in the table following:

Gas	N_2	H_2	NH_3
Concentration (mol L^{-1})	0.0821	0.0583	0.105

For this system:

$$Q = \frac{[NH_3(g)]^2}{[N_2(g)]\,[H_2(g)]^3} = \frac{0.105^2}{0.0821 \times (0.0583)^3}$$

$$= 678$$

Since $Q > K_c$, the reaction will move in the reverse direction.

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Changing the position of equilibrium

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1. The reaction shown in the equation is an example of chemical equilibrium that can be achieved in a beaker – the colour of each species is shown under the formula of the species:

$$\underset{\text{orange}}{Fe^{3+}(aq)} + \underset{\text{colourless}}{SCN^{-}(aq)} \rightleftharpoons \underset{\text{deep blood-red}}{[FeSCN]^{2+}(aq)}$$

Note: The colour of $[FeSCN]^{2+}(aq)$ is so intense that even at very low concentrations, it dominates the appearance of the solution.

Describe and explain your expected observation when the changes **a.**–**c.** are made to a test tube containing these species ($Fe^{3+}(aq)$, $SCN^{-}(aq)$, $[FeSCN]^{2+}(aq)$) in equilibrium.

a. Water is added to the beaker.

b. A few grains of solid sodium thiocyanate, NaSCN, are added to the test tube – the solid dissolves (and thus provides $Na^{+}(aq)$ and $SCN^{-}(aq)$ ions).

c. A few grains of solid sodium fluoride, NaF, are added to the test tube – the solid dissolves (and thus provides $Na^{+}(aq)$ and $F^{-}(aq)$ ions). Fluoride ions combine with iron(III) ions to form a colourless iron-fluoro complex ion.

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d. $SO_3^{2-}(aq) + H_3O^+(aq) \rightleftharpoons HSO_3^-(aq) + H_2O(\ell)$. The equation is balanced.

e. $HSO_3^-(aq) + H_2O(\ell) \rightleftharpoons H_2SO_3(aq) + OH^-(aq)$. Water is an acid in this process.

f. $2H_2O(\ell) \rightleftharpoons H_3O^+(aq) + OH^-(aq)$. One water molecule is acting as a proton donor in this process.

g. $2H_2O(\ell) \rightleftharpoons H_3O^+(aq) + OH^-(aq)$. Neither water molecule is acting as a proton acceptor in this process.

2. a. $SO_4^{2-}(aq) + H_2O(\ell) \rightleftharpoons HSO_4^-(aq) + OH^-(aq)$

b. $HSO_4^-(aq) + H_2O(\ell) \rightleftharpoons SO_4^{2-}(aq) + H_3O^+(aq)$

c. $H_2SO_4(aq) + H_2O(\ell) \rightleftharpoons HSO_4^-(aq) + H_3O^+(aq)$

d. $NH_4^+(aq) + H_2O(\ell) \rightleftharpoons NH_3(aq) + H_3O^+(aq)$

e. $NH_4^+(aq) + OH^-(aq) \rightleftharpoons NH_3(aq) + H_2O(\ell)$

Referring to equations a. to e., identify the ten conjugate acid-base pairs in the five equilibria (some conjugate acid-base pairs appear more than once).

Equation	Conjugate acid	Conjugate base
a.		
a.		
b.		
b.		
c.		
c.		
d.		
d.		
e.		
e.		

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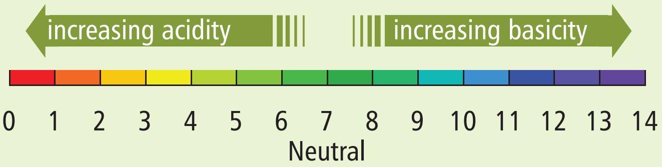
Features of the pH scale

The **pH scale** usually ranges between 0 and 14:

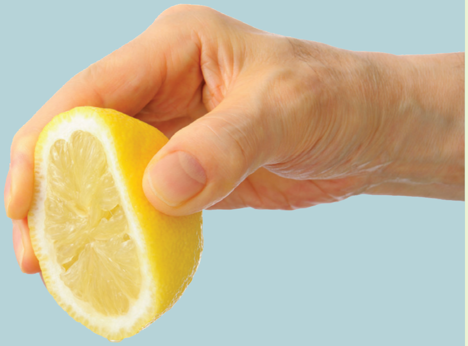
- an acidic solution has a pH number *less* than 7
- a neutral solution has a pH number of *exactly* 7 at 25 °C
- a basic solution has a pH number *greater* than 7

The pH scale



Examples

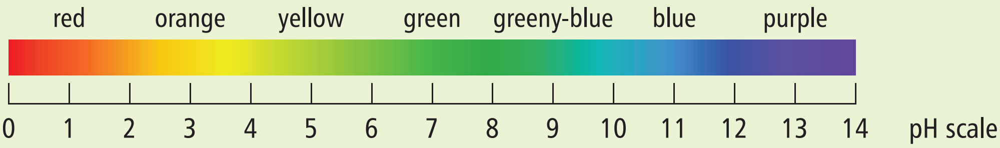
pH of 'everyday' substances	
Substance	**pH**
0.1 mol L^{-1} HCl	1.0
lemon juice	2.5
vinegar	3.0
milk	6.8
pure water	7.0
blood	7.4
window cleaner	11.0
0.1 mol L^{-1} NaOH	13.0



Measurement of pH values

Universal indicator has a range of colours depending upon the pH of a solution.

Universal indicator produces a series of colour changes across the whole pH range



A more accurate value of the pH of a solution is obtained by using a **pH meter**.

Calculating pH

pH can be calculated from the concentration of hydronium (or hydrogen) ions.

Mathematically, pH is defined using logarithms:

pH = negative logarithm of the hydronium ion concentration

$$\text{pH} = -\log\,[H_3O^+(aq)] \qquad (\textit{or}\ \ \text{pH} = -\log_{10}\,[H_3O^+(aq)])$$

Example: Calculating pH

Calculate the pH of a solution where $[H_3O^+] = 4.0 \times 10^{-5}$ mol L^{-1}

$$\text{pH} = -\log\,[H_3O^+]$$
$$= -\log 4.0 \times 10^{-5} \quad \text{substituting}$$
$$= 4.40$$

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Answers

Achievement Standard 91161 (Chemistry 2.1)

Molar mass calculations (page 4)

1. $M(N_2) = 14.0 \times 2 = 28.0$ g mol^{-1}
2. $M(S_8) = 32.0 \times 8 = 256$ g mol^{-1}
3. $M(NaOH) = 23.0 + 16.0 + 1.00 = 40.0$ g mol^{-1}
4. $M(H_3PO_4) = (1.0 \times 3) + 31.0 + (16.0 \times 4) = 98.0$ g mol^{-1}
5. $M(C_6H_{12}O_6) = (12.0 \times 6) + (1.0 \times 12) + (16.0 \times 6) = 180$ g mol^{-1}
6. $M(C_2H_5OH) = (2 \times 12.0) + (6 \times 1.00) + 16.0 = 46$ g mol^{-1}
7. $M(MgCl_2) = 24.3 + (35.5 \times 2) = 95.3$ g mol^{-1}
8. $M(Mg(OH)_2) = 24.3 + (2 \times 16.0) + (2 \times 1.00) = 58.3$ g mol^{-1}
9. $M(Fe_2(SO_4)_3) = (55.9 \times 2) + [3 \times (32.0 + 16.0 \times 4)]$
   $= 111.8 + 288.0 = 400$ g mol^{-1}
10. $M(CuSO_4{\cdot}5H_2O) = 63.5 + 32.0 + 4 \times 16.0 + 5 \times 18.0$
    $= 250$ g mol^{-1}

Mole calculations (page 5)

1. a. i. $m(CO_2) = 1.00$ mol $\times$ 44.0 g mol^{-1} = 44.0 g
   ii. $m(CO_2) = 1.25$ mol $\times$ 44.0 g mol^{-1} = 55.0 g
   iii. $m(CO_2) = 30.0$ mol $\times$ 44.0 g mol^{-1} = 1 320 g
   b. $m(Al_2O_3) = 10.7$ mol $\times$ 102 g mol^{-1} = 1 090 g
   c. $m(SiO_2) = 0.875$ mol $\times$ 60.1 g mol^{-1} = 52.6 g
2. a. $n(CaO) = \dfrac{154 \text{ g}}{56.1 \text{ g mol}^{-1}} = 2.75$ mol
   b. $n(KCl) = \dfrac{8.65 \text{ g}}{74.6 \text{ g mol}^{-1}} = 0.116$ mol
   c. $n(CaCO_3) = \dfrac{4.86 \text{ g}}{100 \text{ g mol}^{-1}} = 0.0486$ mol
   d. $n(NH_4Cl) = \dfrac{2.39 \text{ g}}{53.5 \text{ g mol}^{-1}} = 0.0447$ mol
   e. $n(CuSO_4{\cdot}5H_2O) = \dfrac{5.44 \text{ g}}{250 \text{ g mol}^{-1}} = 0.0218$ mol
3. a. $n(Cl^-) = 0.300$ mol
   b. $n(Cl^-) = 0.500$ mol
   c. $n(NaCl) = \dfrac{4.36 \text{ g}}{58.5 \text{ g mol}^{-1}} = 0.0745$ mol, $n(Cl^-) = 0.0745$ mol
   d. $n(MgCl_2) = \dfrac{140 \text{ g}}{95.3 \text{ g mol}^{-1}} = 1.47$ mol, $n(Cl^-) = 2 \times 1.47 = 2.94$ mol
   e. $n(FeCl_3) = \dfrac{0.372 \text{ g}}{162 \text{ g mol}^{-1}} = 0.00230$ mol, $n(Cl^-) = 3 \times 0.00230$
   $= 0.00690$ mol
4. a. $M(X) = \dfrac{33.5 \text{ g}}{1.38 \text{ mol}} = 24.3$ g mol^{-1}
   b. $M(Y_4) = \dfrac{50.6 \text{ g}}{0.408 \text{ mol}} = 124$ g mol^{-1}, $M(Y) = \dfrac{124}{4} = 31.0$ g mol^{-1}

Reacting masses (page 9)

All calculations should show units in calculation steps.

1. $n(CuO) = \dfrac{1.27 \text{ g}}{79.6 \text{ g mol}^{-1}} = 0.0160$ mol, $n(CuO) : n(Cu) = 1 : 1$
   $n(Cu) = 0.0160$ mol, $m(Cu) = 0.0160$ mol $\times$ 63.6 g mol^{-1} = 1.02 g
2. $n(Mg) = \dfrac{1.68 \text{ g}}{24.3 \text{ g mol}^{-1}} = 0.0691$ mol, $n(Mg) : n(MgO) = 1 : 1$
   $n(MgO) = 0.0691$ mol, $m(MgO) = 0.0691$ mol $\times$ 40.3 g mol^{-1} = 2.78 g
3. $n(Na_2CO_3) = \dfrac{4.88 \text{ g}}{106 \text{ g mol}^{-1}} = 0.0460$ mol, $n(Na_2CO_3) : n(NaHCO_3) = 1 : 2$
   $n(NaHCO_3) = 2 \times 0.0460 = 0.0920$ mol
   $m(NaHCO_3) = 0.0920$ mol $\times$ 84.0 g mol^{-1} = 7.73 g
4. $n(Fe) = \dfrac{2.36 \text{ g}}{55.9 \text{ g mol}^{-1}} = 0.0422$ mol, $n(Fe) : n(FeCl_3) = 1 : 1$
   $n(FeCl_3) = 0.0422$ mol, $m(FeCl_3) = 0.0422$ mol $\times$ 162 g mol^{-1} = 6.84 g
5. a. $n(S) = \dfrac{2.28 \text{ g}}{32.0 \text{ g mol}^{-1}} = 0.0713$ mol, $n(S) : n(FeS) = 1 : 1$, $n(FeS) = 0$
   0713 mol, $m(FeS) = 0.$ 0713 mol $\times$ 87.9 g mol^{-1} = 6.27 g
   b. $n(S) : n(Fe) = 1 : 1$, $n(Fe) = 0.$ 0713 mol, $m(Fe)$ for the reaction to occur = 0. 0713 mol $\times$ 55.9 g mol^{-1} = 4.00 g
   excess $m(Fe)$ = 5.74 g – 4.00 g = 1.74 g
   1.76 using unrounded figures.

Concentration of solutions (page 11)

1. a. i. Concentration (g L^{-1}) = 4.00 g $\times \dfrac{1}{0.100 \text{ L}}$
   = 40.0 g L^{-1}
   ii. $M(NaOH) = 23.0 + 16.0 + 1.0 = 40.0$ g mol^{-1}
   $n(NaOH) = \dfrac{4.00 \text{ g}}{40.0 \text{ g mol}^{-1}}$
   = 0.100 mol
   $V = 0.100$ L
   $c = \dfrac{n}{V} = \dfrac{0.100 \text{ mol}}{0.100 \text{ L}}$
   = 1.00 mol L^{-1}
   b. i. Concentration (g L^{-1}) = 4.00 $\times \dfrac{1}{0.020}$
   = 20.0 g L^{-1}
   ii. $M(KOH) = 39.1 + 16.0 + 1.0 = 56.1$ g mol^{-1}
   $n(KOH) = \dfrac{4.00 \text{ g}}{56.1 \text{ g mol}^{-1}}$
   = 0.0713 mol
   $c = \dfrac{n}{V} = \dfrac{0.0713 \text{ mol}}{0.200 \text{ L}} = 0.357$ mol L^{-1}
   c. i. Concentration (g L^{-1}) = 2.00 g $\times \dfrac{1}{0.200 \text{ L}}$
   = 10.0 g L^{-1}
   ii. $M(H_2C_2O_4) = [2 \times (12.0 + 2 \times 16.0 + 1.0)]$
   $+ [2 \times (2 \times 1.0 + 16.0)]$
   = 126.0 g mol^{-1}
   $n = \dfrac{2.00 \text{ g}}{126.0 \text{ g mol}^{-1}} = 0.0159$ mol
   $c = \dfrac{n}{V} = \dfrac{0.0159 \text{ mol}}{0.200 \text{ L}} = 0.0794$ mol L^{-1}
   d. i. Concentration (g L^{-1}) = 2.12 g $\times \dfrac{1}{0.100 \text{ L}}$
   = 21.2 g L^{-1}
   ii. $M(Na_2CO_3) = (2 \times 23.0) + 12.0 + (3 \times 16.0)$
   = 106.0 g mol^{-1}
   $n = \dfrac{m}{M} = \dfrac{2.12 \text{ g}}{106 \text{ g mol}^{-1}} = 0.0200$ mol
   $c = \dfrac{n}{V} = \dfrac{0.0200 \text{ mol}}{0.100 \text{ L}} = 0.200$ mol L^{-1}
   e. i. Concentration (g L^{-1}) = 0.67 g $\times \dfrac{1}{0.0500 \text{ L}}$
   = 13.4 g L^{-1}

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