1

Fundamentals

Answers to worked examples

WE 1.1 **Derived units (on p. 10 in *Chemistry*³)**

Work out the SI derived unit for molar mass (the mass of one mole of a substance).

Strategy

Write an equation for the molar mass and substitute the SI units for the various quantities.

Solution

Molar mass is the mass of substance per mole.

$$\text{Molar mass} \; = \; \frac{\text{mass}}{\text{number of moles (amount of substance)}}$$

The SI unit of mass is kg and the amount of substance has the unit mol.

$$\text{SI unit of molar mass} \; = \; \frac{\text{SI unit of mass}}{\text{SI unit of amount of substance}}$$

$$= \; \frac{\text{kg}}{\text{mol}} \; = \; \text{kg mol}^{-1}$$

(Note, however, that molar mass (M) is often quoted in units of g mol^{-1}. Be careful to remember to convert molar mass units to kg mol^{-1} before substituting into an equation or calculation.)

WE 1.2 **Atomic symbols (on p. 16 in *Chemistry³*)**

Element X has the atomic symbol $^{234}_{90}X$. What is the name of the element and how many neutrons are there in the nucleus?

Strategy

The superscript on an atomic symbol indicates the atomic mass of the element and the subscript the atomic number. Identify the element from the Periodic Table on the inside front cover.

Solution

The element has atomic number Z = 90, so X is thorium, symbol Th. The atomic number indicates the number of protons, whilst the mass number indicates the number of protons *and* neutrons.

$$\text{Number of neutrons } = \text{ mass number – number of protons } = 234 - 90 = 144$$

WE 1.3 **Relative atomic mass of bromine (on p. 17 in *Chemistry³*)**

Naturally occurring bromine contains ^{79}Br (abundance 50.69%) and ^{81}Br (abundance 49.31%). Calculate the relative atomic mass of naturally occurring bromine.

Strategy

Use the percentage abundances to work out the average relative atomic mass of 100 atoms of bromine. Hence, find the average mass of one atom.

Solution

The average relative atomic mass is given by the weighted average of the masses of the individual isotopes.

Average relative mass of 100 atoms of bromine

$$= (79 \times 50.69) + (81 \times 49.31)$$

$$= 4004.5 + 3994.1$$

$$= 7998.6$$

Average relative mass of one atom of bromine $= A_r(\text{Br}) = 79.99$

(The calculations in this worked example are simplified because isotopic masses are not exactly whole numbers. If you use the precise relative atomic masses for the isotopes in Table 1.5 (p.18), then

$$A_r(Br) = \frac{(50.69 \times 78.9183) + (49.31 \times 80.9163)}{100} = 79.90$$

The value of $A_r(Br)$ is given to four significant figures here because this is the same level of precision as for the relative abundances.)

WE 1.4   Amount of substance (on p. 19 in *Chemistry*³)

What mass of sodium chloride contains 5.82 mol of NaCl?

Strategy

Work out the relative formula mass (M_r) of sodium chloride.

The molar mass of sodium chloride is the relative formula mass in grams.

Use Equation 1.2 to work out the amount in moles.

Solution

$M_r(NaCl) = 22.99 + 35.45 = 58.44$

From Equation 1.2,

$$\text{amount (in mol)} = \frac{\text{mass (in g)}}{\text{molar mass (in g mol}^{-1})}$$

Rearranging the equation,

mass (in g) $= \text{amount (in mol)} \times \text{molar mass (in g mol}^{-1})$

$= 5.82 \text{ mol} \times 58.44 \text{ g mol}^{-1}$

$= 340 \text{ g}$

WE 1.5   Empirical formulae and molecular formulae (on p. 20 in *Chemistry*³)

A compound was found to contain 1.18% hydrogen, 42.00% chlorine, and 56.82% oxygen by mass. What is its empirical formula?

Strategy

Determine the relative amounts, i.e. number of moles of each element in 100 g of the compound, from the relative composition by mass, using Equation 1.2. Find the

simplest ratio of moles of atoms H:Cl:O by dividing by the smallest value. The ratio gives the empirical formula.

Solution

The compound contains 56.82% oxygen by mass. So, 100 g of the compound must contain 56.82 g of oxygen. Using Equation 1.2

amount (in mol) of oxygen in 100 g compound $= 56.82 \text{ g}/16.00 \text{ g mol}^{-1}$

$$= 3.551 \text{ mol}$$

Set out the calculations for the other elements in columns.

	H	**Cl**	**O**
% mass	1.18%	42.00%	56.82%
mass in g	1.18 g	42.00 g	56.82 g
amount (in mol)	$\dfrac{1.18 \text{ g}}{1.01 \text{ g mol}^{-1}}$ $= 1.17$ mol	$\dfrac{42.00 \text{ g}}{35.45 \text{ g mol}^{-1}}$ $= 1.18$ mol	$\dfrac{56.82 \text{ g}}{16.00 \text{ g mol}^{-1}}$ $= 3.55$ mol
simplest ratio (divide by smallest)	$\dfrac{1.17 \text{ mol}}{1.17 \text{ mol}} = 1.00$	$\dfrac{1.18 \text{ mol}}{1.17 \text{ mol}} = 1.01$	$\dfrac{3.55 \text{ mol}}{1.17 \text{ mol}} = 3.00$

Simplest ratio of moles of atoms H:Cl:O is 1:1:3

Empirical formula is $HClO_3$.

WE 1.6   **Using a balanced equation to work out reacting masses (on p. 23 in *Chemistry*[3])**

What mass of oxygen is needed to react exactly with 25 g of ethanol?

Strategy

Write a balanced chemical equation for the reaction of ethanol with oxygen and so determine the stoichiometry of the reaction.

Under the equation, state what the equation tells you about the amounts in moles of ethanol and oxygen that react exactly.

Change the amounts in moles to masses in grams.

Find the mass of oxygen that reacts exactly with 25 g ethanol.

Solution

$$C_2H_5OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 3H_2O(l) \tag{1.3}$$

1 mol   3 mol

46 g 3×32 g $= 96$ g

1 g $(\dfrac{96}{46})$ g

25 g $(\dfrac{96 \times 25}{46})$ g $= 52$ g

So, 52 g of oxygen (O_2) are needed to react exactly with 25 g of ethanol

WE 1.7   **Percentage yield of a reaction (on p. 24 in *Chemistry*[3])**

In the Haber process, nitrogen reacts with hydrogen to form ammonia:

$$N_2 \text{ (g)} + 3 H_2 \text{ (g)} \rightleftharpoons 2 NH_3 \text{ (g)}$$
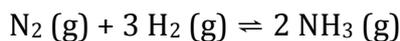
The reaction does not go to completion and an equilibrium mixture of reactants and products is formed. To investigate the equilibrium, 1.00 mol of nitrogen and 3.00 mol of hydrogen are sealed in a container at 100 atm pressure and 400°C. The equilibrium mixture contains 8.57 g of ammonia. What is the percentage yield of the reaction?

Strategy

Use the balanced equation for the reaction to work out the maximum theoretical yield from 8.57 g ammonia. Use Equation 1.4 to calculate the percentage yield from the ratio of the actual to the theoretical yield.

Solution

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

1.00 mol   3.00 mol   2.00 mol

$(2.00 \text{ mol}) \times (17.0 \text{ g mol}^{-1})$
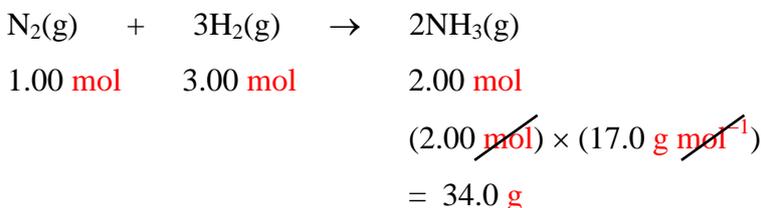
$= 34.0$ g

Using Equation 1.4,

$$\text{percentage yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\% \tag{1.4}$$

$$= \frac{8.57 \text{ g}}{34.0 \text{ g}} \times 100\%$$

$$= 25\%$$

**WE 1.8   Ionic equations (on p. 27 in *Chemistry³*)**

Construct an ionic equation for the reaction of hydrofluoric acid, HF, with a solution of sodium carbonate, Na_2CO_3.
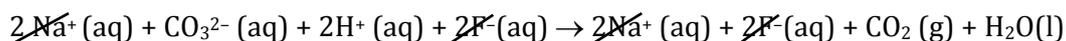
Strategy

Write a full, balanced equation for the reaction, showing the ions present in the reactants and products. Identify any non-ionic substances and include state symbols in the equation. Cross out the spectator ions that appear on both sides of the equation and so do not take part in the reaction.
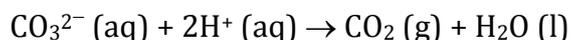
Solution

The full, balanced equation for the reaction is:

$$Na_2CO_3 \text{ (aq)} + 2HF \text{ (aq)} \rightarrow 2NaF \text{ (aq)} + CO_2\text{(g)} + H_2O\text{(l)}$$
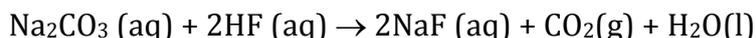
Writing in the ions, and noting that CO_2 and H_2O are covalent compounds,

$$2\cancel{Na^+} \text{ (aq)} + CO_3^{2-} \text{ (aq)} + 2H^+ \text{ (aq)} + \cancel{2F^-}\text{(aq)} \rightarrow 2\cancel{Na^+} \text{ (aq)} + \cancel{2F^-}\text{(aq)} + CO_2 \text{ (g)} + H_2O\text{(l)}$$

and crossing out the spectator ions, the equation becomes:

$$CO_3^{2-} \text{ (aq)} + 2H^+ \text{ (aq)} \rightarrow CO_2 \text{ (g)} + H_2O \text{ (l)}$$

WE 1.9   Naming inorganic compounds (on p. 30 in *Chemistry³*)

What are the systematic names for (a) SeF_4 and (b) $Fe(ClO_4)_3$?

Strategy

Follow the rules for naming inorganic compounds on pp.29 – 30.

Solution

(a) Selenium has more than one oxidation state, so the numbers of atoms in the formula need to be identified. Fluorine is more electronegative, so the compound is selenium tetrafluoride.

(b) Iron forms more than one type of ion, so the oxidation state needs to be indicated. The anion is perchlorate, ClO_4^-, so the compound is iron(III) perchlorate or iron(III) chlorate(VII). Using IUPAC nomenclature, it is iron(III) tetraoxidochlorate(1–).

**WE 1.10** **Changes of oxidation state (on p. 31 in *Chemistry*³)**

Use oxidation states to decide what has been oxidized and what reduced in the reaction of sodium with water.

$$2\ Na\ (s) + 2\ H_2O\ (l) \rightarrow 2\ NaOH\ (aq) + H_2\ (g)$$
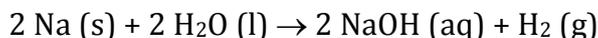
Strategy

Use the rules on p.30-31 to assign oxidation states of each of the element in the equation.

Write the values under the reactants and products in the equation.

An increase in oxidation state on reaction indicates an oxidation process, whilst a decrease indicates a reduction.

Solution
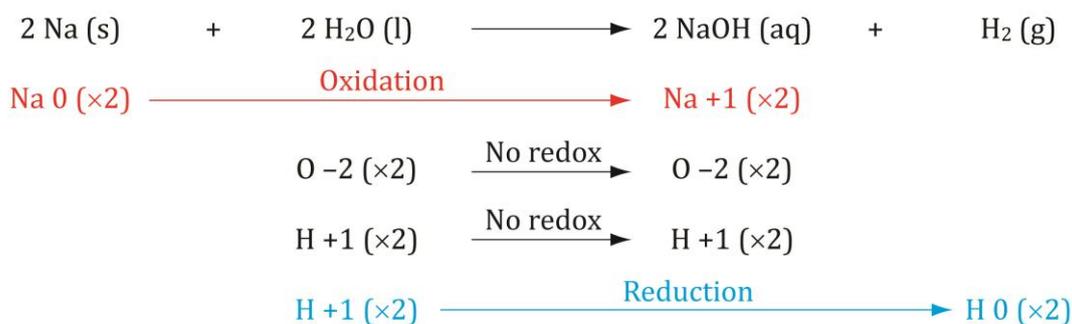
Oxidation state of Na increases from 0 in 2Na to +1 in 2NaOH, so Na has been oxidized.

Oxidation state of H decreases from +1 in $2H_2O$ to 0 in H_2, so H has been reduced.

(Note that there are also two H in $2H_2O$ that end up in 2NaOH. The oxidation state of H in both these is +1, so no redox.)

WE 1.11 **Constructing and balancing a redox equation from half equations (on p. 32 in *Chemistry*³)**

Fe^{2+} ions are oxidized to Fe^{3+} by MnO_4^- ions in acidic solution. The MnO_4^- ions are reduced to Mn^{2+} ions. Construct half equations and a balanced overall equation for the reaction.
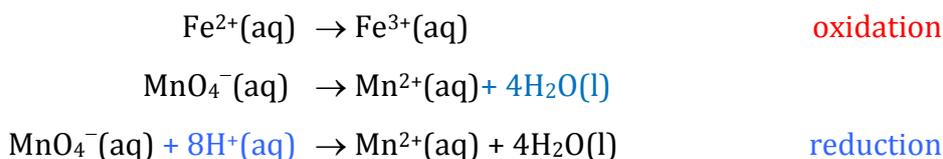
Strategy

Start by writing unbalanced half equations for the oxidation of Fe^{2+} ions and the reduction of MnO_4^- ions. Then follow through the steps 1-8 on p.32.
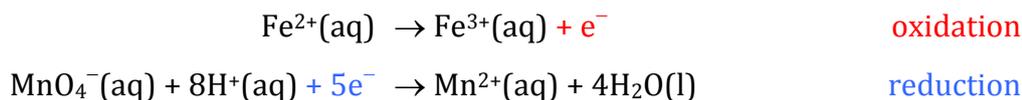
Solution

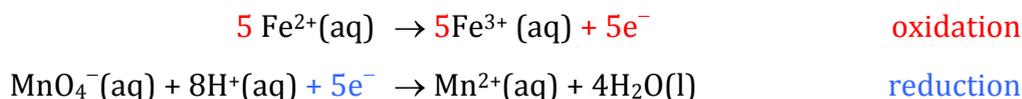**Steps 1and 2:** Write down the unbalanced half equations. Balance elements other than O and H.

$$Fe^{2+} (aq) \rightarrow Fe^{3+} (aq) \qquad \text{oxidation}$$

$$MnO_4^- (aq) \rightarrow Mn^{2+} (aq) \qquad \text{reduction}$$

Step 3: Balance O and H by adding H_2O and H^+ ions (in acidic solution)

$$Fe^{2+}(aq) \rightarrow Fe^{3+}(aq) \qquad \text{oxidation}$$

$$MnO_4^-(aq) \rightarrow Mn^{2+}(aq) + 4H_2O(l)$$

$$MnO_4^-(aq) + 8H^+(aq) \rightarrow Mn^{2+}(aq) + 4H_2O(l) \qquad \text{reduction}$$

Step 4: Balance electric charge by adding electrons

$$Fe^{2+}(aq) \rightarrow Fe^{3+}(aq) + e^- \qquad \text{oxidation}$$

$$MnO_4^-(aq) + 8H^+(aq) + 5e^- \rightarrow Mn^{2+}(aq) + 4H_2O(l) \qquad \text{reduction}$$

Step 5: Multiply oxidation half equation ×5, so number of electrons transferred is same in both half equations.

$$5 Fe^{2+}(aq) \rightarrow 5Fe^{3+} (aq) + 5e^- \qquad \text{oxidation}$$

$$MnO_4^-(aq) + 8H^+(aq) + 5e^- \rightarrow Mn^{2+}(aq) + 4H_2O(l) \qquad \text{reduction}$$

Step 6: Add the two half equations and cancel out the electrons

$$5 Fe^{2+}(aq) \rightarrow 5 Fe^{3+}(aq) + \cancel{5e^-} \qquad \text{oxidation}$$

$$MnO_4^-(aq) + 8 H^+(aq) + \cancel{5e^-} \rightarrow Mn^{2+}(aq) + 4 H_2O(l) \qquad \text{reduction}$$

$$MnO_4^- (aq) + 8 H^+(aq) + 5 Fe^{2+}(aq) \rightarrow 5 Fe^{3+}(aq) + Mn^{2+}(aq) + 4 H_2O(l) \qquad \text{overall}$$
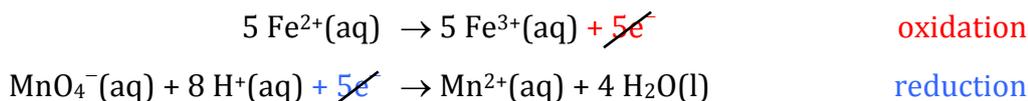
Steps 7 and 8: No further simplification is necessary as atoms and charges balance.

WE 1.12 Units of concentration (on p. 35 in *Chemistry*[3])

43.7 g of anhydrous copper(II) sulfate ($CuSO_4$) were dissolved in water and the solution made up to 2.00 dm^3. What is the molarity of the solution?

Strategy

Work out the concentration of the solution in $g \ dm^{-3}$.

Use Equation 1.8 to convert $g \ dm^{-3}$ to $mol \ dm^{-3}$. Pay particular attention to the units.

Solution

$M_r(CuSO_4) = 63.55 + 32.07 + (4 \times 16.00) = 159.6$

Concentration $(g\ dm^{-3}) = \dfrac{43.7\ g}{2.00\ dm^3} = 21.85\ g\ dm^{-3}$

Using Equation 1.8,

Concentration $(mol\ dm^{-3}) = \dfrac{\text{concentration } (g\ dm^{-3})}{\text{molar mass } (g\ mol^{-1})}$ $\hspace{2cm}$ (1.8)

$\hspace{3cm} = \dfrac{21.85\ g\ dm^{-3}}{159.6\ g\ mol^{-1}}$

$\hspace{3cm} = 0.137\ mol\ dm^{-3}$

WE 1.13 **Finding out how much solute is in a given volume of solution (on p. 36 in *Chemistry*³)**

(a) A solution of sodium carbonate has a concentration of 0.157 mol dm⁻³. What volume of solution contains exactly 1.00 g of sodium carbonate?

(b) What mass of sodium thiosulfate crystals ($Na_2S_2O_3.5H_2O$) must be dissolved in 250 cm³ of water to give a solution with a concentration of 0.0750 mol dm⁻³?

(a) **Strategy**

Use Equation 1.2 to find the amount (in mol) of sodium carbonate in 1.00 g. Use the definition of concentration to determine the volume of solution. You can do this from first principles or by using Equation 1.9 (but remember to convert the volume of the solution to dm³ (1 dm³ = 1000 cm³)).

Solution

$M_r(Na_2CO_3) = (2 \times 22.99) + 12.00 + (3 \times 16.00) = 106.0$

Using Equation 1.2,

amount of Na_2CO_3 in 1.00 g $= \dfrac{1.00\ g}{106.0\ g\ mol^{-1}} = 9.43 \times 10^{-3}\ mol$

Working out the volume from first principles.

0.157 mol are contained in 1.00 dm³ of solution

1 mol is contained in $\dfrac{1.00}{0.157}$ dm³ of solution

9.43×10^{-3} mol are contained in $\dfrac{1.00 \times (9.43 \times 10^{-3})}{0.157}$ dm^3 of solution

$$= 0.0601 \text{ dm}^3 \text{ of solution}$$

$$= 60.1 \text{ cm}^3$$

Working out the volume using Equation 1.9 on p.36

Using Equation 1.9, $n = c \times V$

$$V = \frac{n}{c} = \frac{9.43 \times 10^{-3} \text{ mol}}{0.157 \text{ mol dm}^{-3}}$$

$$= 0.0601 \text{ dm}^3$$

$$= 60.1 \text{ cm}^3$$

(b)   **Strategy**

Calculate the amount in mol of sodium thiosulfate in 250 cm^3 of solution.
Work out the molar mass of $Na_2S_2O_3.5H_2O$ and rearrange Equation 1.2 to
convert the amount in mol to the mass of crystals needed.

Solution

$M_r(Na_2S_2O_3.5H_2O) = (2 \times 22.99) + (2 \times 32.07) + (3 \times 16.00) + (5 \times 18.02) = 248.2$

1000 cm^3 of solution contain 0.0750 mol of sodium thiosulfate

250 cm^3 of solution contain 0.0750/4 mol $= 0.01875$ mol of sodium thiosulfate

Mass of $Na_2S_2O_3.5H_2O$ needed $= 248.2$ g mol^{-1} $\times 0.01875$ mol $= 4.65$ g

WE 1.14   An acid-base titration (on p. 37 in *Chemistry*³)

14.8 cm³ of 0.00105 mol dm⁻³ nitric acid, HNO_3 (aq), reacted exactly with 9.85 cm³ of
calcium hydroxide solution, $Ca(OH)_2$ (aq). Calculate the concentration of the calcium
hydroxide solution in (a) mol dm⁻³ and (b) in g dm⁻³.

Strategy

Working through steps 1-5 on p.37:

write a balanced equation for the reaction;

state what the equation tells you about the amounts of HNO_3 (aq) and $Ca(OH)_2$ (aq)
that react exactly;

use the concentration of the standard solution of nitric acid to work out the number
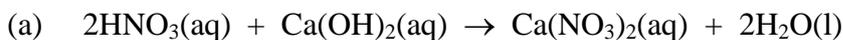of moles of HNO_3 that have reacted;

use the balanced equation to work out the amount (in mol) of $Ca(OH)_2$ that reacted

exactly with this amount of HNO_3

convert this to the amount in 1 dm^3 to give the answer in $mol\ dm^{-3}$;

Work out the molar mass of $Ca(OH)_2$ and use Equation 1.8 on p.35 to find the concentration in $g\ dm^{-3}$.

Solution

(a) $2HNO_3(aq)\ +\ Ca(OH)_2(aq)\ \rightarrow\ Ca(NO_3)_2(aq)\ +\ 2H_2O(l)$

     2 mol   1 mol

     14.8 cm^3 of 0.00105 $mol\ dm^{-3}$ $HNO_3(aq)$ were added

     Working out the amount of HNO_3 added from first principles

     1000 cm^3 of $HNO_3(aq)$ contain 0.00105 mol of HNO_3

     1 cm^3 of $HNO_3(aq)$ contains $\dfrac{0.00105}{1000}$ mol of HNO_3

     14.8 cm^3 of $HNO_3(aq)$ contain $\dfrac{0.00105 \times 14.8}{1000}$ mol $=\ 1.554 \times 10^{-5}$ mol of HNO_3

     Working out the amount of HNO_3 added using Equation 1.9 on p.36

     $n\ =\ c \times V$ (1.9)

     Amount of HNO_3 added $=\ (0.00105\ mol\ dm^{-3}) \times (0.0148\ dm^3)$

                                       $=\ 1.554 \times 10^{-5}$ mol of HNO_3

     (Remember to convert the volume of the solution to dm^3)

     Amount of $Ca(OH)_2$ to react exactly $=\ 0.5 \times (1.554 \times 10^{-5}\ mol)$

     This was contained in 9.85 cm^3 of $Ca(OH)_2(aq)$

     Concentration of $Ca(OH)_2(aq) = \dfrac{0.5 \times (1.554 \times 10^{-5}\ mol)}{0.00985\ dm^3}$

                                      $=\ 7.89 \times 10^{-4}\ mol\ dm^{-3}$ $(=\ 0.000789\ mol\ dm^{-3})$

(b) $M_r(Ca(OH)_2)\ =\ 40.08\ +\ (2 \times 16.00)\ +\ (2 \times 1.01)\ =\ 74.10$

     From Equation 1.8 on p.35,

     Concentration of $Ca(OH)_2(aq)$ in $g\ dm^{-3}$

                           $=\ 7.89 \times 10^{-4}\ mol\ dm^{-3} \times 74.10\ g\ mol^{-1}$

                           $=\ 0.0585\ g\ dm^{-3}$

WE 1.15   Gravimetric analysis (on p. 40 in *Chemistry³*)

A 4.500 g sample of an oil containing the pesticide DDT ($C_{14}H_9Cl_5$) was heated with sodium in alcohol to liberate all the chlorine present as Cl^- ions. The mixture was treated with silver nitrate solution and 1.509 g of AgCl were recovered. Calculate the percentage (by mass) of DDT in the sample.

<u>Strategy</u>

Work out the amount (in mol) of AgCl recovered.

Use the stoichiometry of the reaction of DDT with sodium to determine the amount of DDT in the sample of oil.

Calculate the molar mass of DDT and thus the total mass of DDT present in the sample of oil.

Finally, find the percentage by mass as the ratio of the total mass of DDT to the mass of the oil sample.

<u>Solution</u>

$M_r(AgCl) = 107.87 + 35.45 = 143.3$

Mass of AgCl formed $= 1.509$ g

Amount of AgCl formed $= \dfrac{1.509 \text{ g}}{143.3 \text{ g mol}^{-1}}$

$= 1.053 \times 10^{-2}$ mol

1 mol of DDT ($C_{14}H_9Cl_5$) reacts to form 5 mol Cl^- ions

So, 4.500 g of the oil contain $\dfrac{1.053 \times 10^{-2} \text{ mol}}{5}$ of DDT

$= 2.1061 \times 10^{-3}$ mol of DDT

$M_r(C_{14}H_9Cl_5) = (14 \times 12.01) + (9 \times 1.01) + (5 \times 35.45) = 354.5$

Mass of DDT $= (2.1061 \times 10^{-3} \text{ mol}) \times (354.5 \text{ g mol}^{-1})$

$= 0.7466$ g

% of DDT in the oil $= \dfrac{0.7466 \text{ g}}{4.500 \text{ g}} \times 100$

$= 16.6\%$

WE 1.16 Using thermochemical equations (on p. 44 in *Chemistry³*)

When 1 mol of pentane burns in a plentiful supply of air, 3537 kJ of energy are transferred as heat to the surroundings at 298 K. Write a balanced thermochemical equation for the reaction, and use it to work out the enthalpy change when 500 g of pentane are burned.

Strategy

Write a balanced thermochemical equation for the reaction of pentane with oxygen, showing the enthalpy change when 1 mol of pentane reacts.

Use Equation 1.2 on p.19 to work out the amount in moles contained in 500 g of pentane.

Find the enthalpy change when this amount of pentane is burned.

Solution

$$C_5H_{12}(l) \; + \; 8O_2(g) \; \rightarrow \; 5CO_2(g) \; + \; 6H_2O(l) \qquad\qquad \Delta_rH^{\ominus}_{298} \; = \; -3537 \text{ kJ mol}^{-1}$$

$M_r(C_5H_{12}) \; = \; 72.1$

500 g of pentane contain $\dfrac{500\,\cancel{g}}{72.1\,\cancel{g}\text{ mol}^{-1}} \; = \; 6.94$ mol

Enthalpy change on burning 500 g pentane $\quad = \; -3537 \text{ kJ } \cancel{\text{mol}^{-1}} \times 6.94 \; \cancel{\text{mol}}$
$\qquad\qquad\qquad\qquad\qquad\qquad\qquad\qquad = \; -24\,547 \text{ kJ} \approx -24\,500 \text{ kJ}$

WE 1.17 Equilibrium constants in terms of concentrations (on p. 61 in *Chemistry³*)

An equimolar mixture of ethanoic acid and ethanol was heated at 373 K. At equilibrium, the concentration of ethanoic acid was found to be 0.820 mol dm⁻³. Calculate the concentration of ethyl ethanoate in the equilibrium mixture. (*Hint*. Use the stoichiometric equation for the reaction to determine the concentration of ethanol in the equilibrium mixture.)

Strategy

Write a balanced chemical equation for the reaction and use it to produce an expression for the equilibrium constant in terms of the concentrations of the products and reactants. Rearrange the expression, using the value for the equilibrium constant in Worked Example 1.17 to determine the concentration of ethyl ethanoate.

Solution

The balanced equation for the reaction between ethanoic acid and ethanol is

At 373 K, $\quad K_c \quad = 4.00$ (see worked example)

$$= \frac{[CH_3CO_2C_2H_5(l)][H_2O(l)]}{[CH_3CO_2H(l)][C_2H_5OH(l)]}$$

An equimolar mixture of CH_3CO_2H and C_2H_5OH was heated, so at equilibrium

$$[CH_3CO_2C_2H_5(l)] \quad = [H_2O(l)]$$

and $\qquad [CH_3CO_2H(l)] \qquad = [C_2H_5OH(l)]$

The expression for the equilibrium constant can now be written as

$$K_c = \frac{[CH_3CO_2C_2H_5(l)]^2}{(0.820 \text{ mol dm}^{-3})^2} = 4.00$$

Taking the square root and rearranging

$$[CH_3CO_2C_2H_5(l)] \qquad = 2.00 \times 0.820 \text{ mol dm}^{-3}$$

$$= 1.64 \text{ mol dm}^{-3}$$

WE 1.18 **Equilibrium constants in terms of partial pressures (on p. 63 in *Chemistry³*)**

A second investigation was carried out at the same temperature (673 K) but at a higher pressure. At equilibrium, the partial pressure of N_2 was 18.7 atm and the partial pressure of hydrogen was 56.1 atm.

(a) Calculate the partial pressure of NH_3 in the equilibrium mixture.

(b) What was (i) the total pressure for the second investigation and (ii) the percentage of NH_3 in the equilibrium mixture?

Strategy

(a) Use the expression for the equilibrium constant and its value at 673 K from Worked Example 1.18 and substitute the values for the partial pressures of H_2 and N_2 in the second investigation. Rearrange the equation to find the partial pressure of NH_3 in the equilibrium mixture.

(b) Calculate the total pressure by summing the partial pressures for each of the components. Determine the percentage of NH_3 in the reaction mixture from the ratio of the partial pressure of NH_3 to the total pressure.

Solution

(a) At 673 K, $\qquad K_p \quad = 1.7 \times 10^{-4} \text{ atm}^{-2}$ (see worked example)

$$= \frac{p_{NH_3}^2}{p_{N_2}\, p_{H_2}^3}$$

$$= \frac{p_{NH_3}^2}{18.7 \text{ atm} \times (56.1 \text{ atm})^3}$$

$$p_{NH_3}^2 = (1.7 \times 10^{-4} \text{ atm}^{-2}) \times (18.7 \text{ atm}) \times (56.1 \text{ atm})^3$$

$$= 561.3 \text{ atm}^2$$

$$p_{NH_3} = 23.7 \text{ atm}$$

(b) (i) $p_{total} = p_{N_2} + p_{H_2} + p_{NH_3}$

$$= 18.7 \text{ atm} + 56.1 \text{ atm} + 23.7 \text{ atm}$$

$$= 98.5 \text{ atm}$$

(ii) Percentage of NH_3 in equilibrium mixture $= \dfrac{23.7}{98.5} \times 100 = 24\%$

Answers to questions in boxes

Box 1.2 **Mass spectrometry (on p. 14 in *Chemistry*³)**

Look at the mass spectrum of magnesium in Figure 2(b). How many protons and neutrons are there in the nuclei of each of the three isotopes?

Strategy

Identify the isotopes from the mass spectrum. The atomic number gives the number of protons. Deduce the number of neutrons from the difference between the mass number and the atomic number.

Solution

The mass spectrum contains the ions ^{24}Mg, ^{25}Mg and ^{26}Mg. The superscript indicates the mass number.

All have 12 protons, as the atomic number $Z = 12$ for magnesium.

^{24}Mg has 12 neutrons, ^{25}Mg has 13 neutrons and ^{26}Mg has 14 neutrons.

Box 1.3 **The Amesbury Archer (on p. 17 in *Chemistry*³)**

How do the three isotopes of oxygen vary in the number of protons, neutrons, and electrons present?

Strategy

The three isotopes of oxygen vary due to the number of neutrons in the nucleus. The atomic number gives the number of protons (and electrons) in each isotope. Work out the number of neutrons in each isotope from the difference between the mass number and the atomic number.

Solution

The atomic number of oxygen is 8, so that all isotopes have 8 protons (and 8 electrons in the neutral atoms).

The number of neutrons is given by the difference between the mass number and the atomic number, so for ^{16}O it is 8, for ^{17}O it is 9, and for ^{18}O it is 10.

Box 1.4 **Atom efficiency and green chemistry (on p. 25 in *Chemistry*³)**

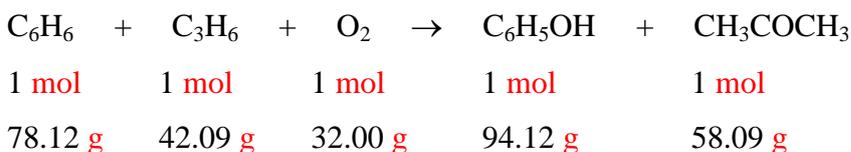Demand for propanone is rising at a lower rate than the demand for phenol.

Calculate the atom efficiency of the cumene process if half the propanone is considered as waste and the overall conversion to products is 90%.

Strategy

Use the overall equation for the three stage process given in Box 1.4 on p.25. Under the equation, write the stoichiometry (for 1 mol of each reactant) and the reacting masses of the reactant and products

Use Equation 1.5 to determine the atom efficiency, remembering that only half the mass of propanone is useful.

Solution

$$C_6H_6 \ + \ C_3H_6 \ + \ O_2 \ \rightarrow \ C_6H_5OH \ + \ CH_3COCH_3$$

1 mol   1 mol   1 mol   1 mol   1 mol

78.12 g   42.09 g   32.00 g   94.12 g   58.09 g

$M_r (C_6H_6) = [(6 \times 12.01)] + (6 \times 1.01)] = 78.12$

$M_r (C_3H_6) = [(3 \times 12.01)] + (6 \times 1.01)] = 42.09$

$M_r (O_2) = [(2 \times 16.00)] = 32.00$

$M_r (C_6H_5OH) = [(6 \times 12.01)] + (6 \times 1.01) + 16.00] = 94.12$

$M_r (CH_3COCH_3) = [(3 \times 12.01)] + (6 \times 1.01) + 16.00] = 58.09)$

If half the propanone is considered to be waste, only $\dfrac{58.09 \text{ g}}{2} = 29.05$ g count as desired product.

Using Equation 1.5 and assuming 100% conversion to products,

$$\text{atom efficiency} \quad = \frac{94.12 \text{ g mol}^{-1} + 29.05 \text{ g mol}^{-1}}{78.12 \text{ g mol}^{-1} + 42.09 \text{ g mol}^{-1} + 32.00 \text{ g mol}^{-1}} \times 100\%$$

$$= \frac{123.17 \text{ g mol}^{-1}}{152.21 \text{ g mol}^{-1}} \times 100\% = 80.9\%$$

For 90% conversion to products,

$$\text{atom efficiency} \quad = \frac{(0.9 \times 123.17) \text{ g}}{152.21 \text{ g}} \times 100\% = 72.8\%$$

Box 1.5   Measuring low concentrations (on p. 34 in *Chemistry*³)

The label on a bottle of mineral water says it contains '27 mg calcium per litre'.

(a) What is the concentration of Ca^{2+} ions in ppm (by mass) and the percentage by mass of Ca^{2+} ions?

(b) What is the concentration of Ca^{2+} ions in mol dm^{-3}?

Strategy

Assume that the density of the mineral water is 1.0 g dm^{-3} and work out the mass of one litre. Using Equation 1.6, calculate the proportion of Ca^{2+} ions by mass as a fraction of the total mass.

Solution

(a) Assuming that the density of mineral water is 1.0 g cm^{-3}, one litre of water, which is equivalent to a volume of 1×10^3 cm^3, has, from Worked Example 1.1, a mass of

$$m = \rho V = 1.0 \text{ g cm}^{-3} \times (1 \times 10^3) \text{cm}^3 = 1.0 \times 10^3 \text{g} = 1.0 \text{ kg}$$

Then, from Equation 1.6, the concentration by mass of Ca^{2+} ions is

$$\frac{m_{Ca^{2+}}}{m_{total}} \approx \frac{m_{Ca^{2+}}}{m_{H_2O}}$$

$$= \frac{0.027 \text{ g}}{1000 \text{ g}} \times 10^6 \text{ ppm} = 27 \text{ ppm}$$

$$\% \text{ Composition} = \frac{0.027 \text{ g}}{1000 \text{ g}} \times 100\% = 0.0027\%$$

(b) $A_r(Ca) = 40.1$
From Equation 1.8,

$$\text{concentration (mol dm}^{-3}) = \frac{27 \times 10^{-3} \text{ g dm}^{-3}}{40.1 \text{ g mol}^{-1}} = 6.7 \times 10^{-4} \text{ mol dm}^{-3}$$

Box 1.6 **Measuring dissolved oxygen in river water (on p. 38 in *Chemistry*[3])**

A Winkler titration was performed on a 150 cm^3 sample of river water. 12.2 cm^3 of a 0.0100 mol dm^{-3} sodium thiosulfate solution were required.

(a) Calculate the concentration of the dissolved oxygen in mg dm^{-3}.

(b) Suggest why a back titration is used in this determination, rather than a direct titration of O_2 against Mn^{2+} (aq).

Strategy

Determine the amount (in mol) of thiosulfate, $S_2O_3^{2-}$(aq) used in the titration. (You can do this from first principles, or by using Equation 1.9 but in this case remember to convert the volume of the solution to dm^3.)

Use Equation 1.12 to find the amount of iodine, I_2 (aq), that must have reacted with $S_2O_3^{2-}$(aq) in the back titration.

Use Equations 1.10, 1.11 and 1.12 to calculate the amount of oxygen, O_2, that must have been present in the 150 cm^3 sample to generate this amount of iodine, and hence the concentration of oxygen in the river water.

Solution

(a)   *Working out the amount of $S_2O_3^{2-}$ ions added from first principles*

1000 cm^3 of $Na_2S_2O_3$(aq) contain 0.0100 mol of $S_2O_3^{2-}$ ions

1 cm^3 of $Na_2S_2O_3$(aq) contains $\dfrac{0.0100}{1000}$ mol of $S_2O_3^{2-}$ ions

12.2 cm^3 of $Na_2S_2O_3$(aq) contain $\dfrac{0.0100 \times 12.2}{1000}$ mol $= 1.22 \times 10^{-4}$ mol of $S_2O_3^{2-}$ ions

Working out the amount of $S_2O_3^{2-}$ ions added using Equation 1.9 on p.36

$$n = c \times V \qquad\qquad (1.9)$$

Amount of $S_2O_3^{2-}$ ions added $\quad = 0.0100$ mol $\cancel{dm^{-3}} \times 0.0122 \cancel{\,dm^3}$

$\qquad\qquad = 1.22 \times 10^{-4}$ mol

From Equation 1.12, amount of I_2(aq) that reacted with the $S_2O_3^{2-}$ ions in the back titration $\qquad\qquad = 0.5 \times (1.22 \times 10^{-4}$ mol$)$

From Equations 1.10, 1.11 and 1.12, 1 mol of I_2 is generated from 0.5 mol of O_2(aq).
Amount of O_2(aq) in the 150 cm^3 sample of river water

$\qquad\qquad = 0.5 \times 0.5 \times 1.22 \times 10^{-4}$ mol

$\qquad\qquad = 3.05 \times 10^{-5}$ mol

Concentration of O_2(aq) in mol dm^{-3} $\quad = (3.05 \times 10^{-5} \times \dfrac{1000}{150})$ mol dm^{-3}

$\qquad\qquad = 2.033 \times 10^{-4}$ mol dm^{-3}

$A_r(O_2) = 32.00$

From Equation 1.8 on p.35,

Concentration of O_2(aq) in g dm^{-3} $\quad = (2.033 \times 10^{-4} \cancel{\text{ mol } dm^{-3}}) \times 32.00$ g $\cancel{mol^{-1}}$

$$= 6.51 \times 10^{-3} \text{ g dm}^{-3}$$

$$= 6.51 \text{ mg dm}^{-3}$$

(b) For an accurate determination, all the dissolved oxygen must react with the $Mn(OH)_2(s)$, so it is important that the reaction takes place quickly. Furthermore, the end point of a direct titration would be very difficult to determine.

Box 1.7 Potential energy and kinetic energy (on p. 42 in *Chemistry*^{3})

(a) Calculate the energy needed for a bungee jumper weighing 65.0 kg to climb from the ground to a platform 35.0 m high.

(b) Calculate the kinetic energy of a nitrogen molecule travelling at 500 m s^{-1}.

Strategy

Use Equation 1.16, which defines how the potential energy of an object varies with its height above the surface of the Earth, and Equation 1.17, which defines an object's kinetic energy.

Solution

(a) The gravitational potential energy of the person standing on the platform is equal to the energy needed for the person to climb from the ground to the platform. From Equation 1.16,

$$E_{PE} = mgh = 65.0 \text{ kg} \times 9.81 \text{ m s}^{-2} \times 35.0 \text{ m} = 22300 \text{ kg m}^2\text{s}^{-2} = 22.3 \text{ kJ}$$

(Remember 1 J = 1 kg m^2 s^{-2} (Table 1.2))

(b) M_r (N$_2$) $= 2 \times 14.007 = 28.01$

Molar mass (N$_2$), $M = 28.01 \times 10^{-3}$ kg mol^{-1}

Mass of an individual nitrogen molecule, m, is given by

$$m = M/N_A = 28.01 \times 10^{-3} \text{ kg mol}^{-1}/6.022 \times 10^{23} \text{ mol}^{-1}$$

$$= 4.651 \times 10^{-21} \text{ kg}$$

From Equation 1.17,

$$E_{KE} = \tfrac{1}{2}mv^2 = 4.651 \times 10^{-26}\text{kg} \times (500 \text{ m s}^{-1})^2 = 5.81 \times 10^{-21} \text{ J}$$

Box 1.9   Butane hair stylers (on p. 45 in *Chemistry*³)

The label on a 'butane' refill for a hair styler states that it contains 14 g of 'isobutane'. The formula of 'isobutane' is C_4H_{10}.

(a) Assuming the fuel burns completely, according to the thermochemical equation above, calculate the heat energy supplied from 14 g of 'isobutane'.

(b) Suggest a structural formula (see Section 1.3, p.12) for 'isobutane' and give its systematic IUPAC name.

(You can read about structural isomers and naming hydrocarbons in Section 2.5, p.79. You will be able to calculate enthalpy changes under more realistic conditions, not just 25°C, after studying Section 13.4, p.631.)

Strategy

Calculate the amount in moles that is equivalent to 14 g 'isobutane'. The thermochemical equation tells you how much heat is given out by 1 mol, so you can work out how much heat is given out by the amount of 'isobutane' in the question.

Solution

(a)   From the thermochemical equation, 2878 kJ of energy are supplied when 1 mol of butane reacts with oxygen at 298 K
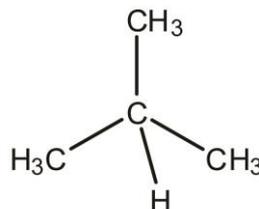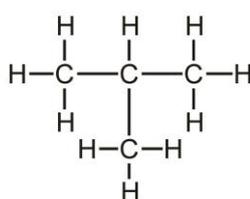
$M_r(C_4H_{10})$ = 58.1

14 g isobutane contain $\dfrac{14\ \cancel{g}}{58.\cancel{1}\ \cancel{g}\ mol^{-1}}$ = 0.241 mol

Heat supplied by burning 14 g isobutane

   = 2878 kJ $\cancel{mol^{-1}}$ × 0.241 $\cancel{mol}$

   = 690 kJ

(b)   Drawing the structure of isobutene in either a two- or three-dimensional form shows that the systematic IUPAC name is 2-methylpropane.

(2-)methylpropane

Box 1.10 Phase changes of water (on p. 50 in *Chemistry³*)

(a) In warm countries, water can be kept cool by storing it in unglazed earthenware pots. Explain the process involved.

(b) Suggest why wet clothes dry on a washing line but not in a plastic bag.

(c) Using the data in Figure 3, write down values at 0°C for: (i) $\Delta_{sub}H^{\ominus}(H_2O)$; (ii) $\Delta_{cond}H^{\ominus}(H_2O)$; (iii) $\Delta_{freezing}H^{\ominus}(H_2O)$.

(d) Suggest why the value of $\Delta_{vap}H^{\ominus}(H_2O)$ is very much more endothermic than the value of $\Delta_{fus}H^{\ominus}(H_2O)$.

Strategy

Think about the equilibria between the different phases of water and the enthalpy changes involved. Pay particular attention to the signs of the enthalpy changes.

Solution

(a) Water seeps into the clay and evaporates slowly from the surface of the pot in the warm climate. Evaporation is an endothermic process and cools the pot.

(b) In a plastic bag, an equilibrium is set up in which the rate of evaporation of water from the wet clothes equals the rate of condensation. On a washing line, no equilibrium can be established. Water vapour is removed in air currents and the equilibrium constantly moves towards the right-hand side: $H_2O(l) \rightleftharpoons H_2O(g)$

(c) (i) +51.1 kJ mol⁻¹   (ii) −4.5 kJ mol⁻¹   (iii) −6.01 kJ mol⁻¹

(d) There is a much larger change in structure from liquid water → water vapour than from ice → liquid water. Hydrogen bonds are present in both ice and liquid water, but all these are broken when the vapour is formed.

Box 1.11 Why is Kevlar so strong? (on p. 55 in *Chemistry³*)

In terms of intermolecular attractions, suggest why it is possible to make strong ropes out of KEVLAR® or nylon, but not from polythene (poly(ethene)).

Strategy

Consider the nature of the interactions between the adjacent chains in the three polymers.

Solution

Hydrogen bonding between adjacent chains in Kevlar and nylon leads to ordered structures in which the aligned chains cannot move past one another very easily. In

poly(ethene), the only intermolecular attractions between the chains are the weaker London dispersion interactions. So, fibres made from nylon or Kevlar are stronger than fibres made from poly(ethene).

**Box 1.12 Connecting equilibria and cave chemistry (on p. 58 in *Chemistry³*)**

(a) Use Le Chatelier's principle and Equation 1.22 to explain both the formation of caves and the presence of stalactites and stalagmites in limestone areas.

(b) The concentration of carbon dioxide in the atmosphere is currently increasing from year to year. What effect does this increase in carbon dioxide concentration have on the acidity of rainwater?

Strategy

(a) When applying Le Chatelier's Principle to the equilibrium in Equation 1.22, bear in mind that the concentration of carbon dioxide in the air seeping through limestone rocks is higher than that in the air inside the cave.

(b) Consider the effect of increasing carbon dioxide concentration on the equilibria in Equations 1.19 and 1.20.

Solution

(a) The erosion of limestone results from the reaction

$$CaCO_3 \text{ (s)} + CO_2 \text{ (g)} + H_2O \text{ (l)} \rightleftharpoons Ca^{2+} \text{ (aq)} + 2HCO_3^- \text{ (aq)}$$
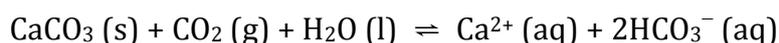
The concentration of carbon dioxide in the air, CO_2 (g), that is in contact with the water seeping through the limestone rocks is higher than the normal atmospheric concentration, so the position of equilibrium moves to the right to minimize the effect of the imposed change. $CaCO_3$ (s) dissolves to form Ca^{2+} (aq) ions and HCO_3^- (aq) ions, creating caves in the rock. When the water drips from the roof of the cave, CO_2 (g) in the surrounding air is the normal atmospheric concentration. The position of equilibrium now moves to the left to minimize the effect of the change and $CaCO_3$ (s) precipitates from the solution forming stalactites and stalagmites.

(b) An increase in CO_2 (g) will result in an increase in CO_2 (aq) (Equation 1.19). This, in turn, will cause the position of equilibrium in Equation 1.20 to move to the

right. The concentration of H$^+$ (aq) will increase and rain water will become more acidic.

Answers to end of chapter questions

1. The C–C bond length in a crystal of diamond is 0.154 nm. What is this distance in (a) metres, (b) picometres, (c) ångstroms? (Section 1.2)

Strategy

Use the SI prefixes in Table 1.3 to convert the bond length. Remember that 1 Å is equivalent to 1×10^{-10} m.

Solution

(a) 1.54×10^{-10} m

(b) 154 pm

(c) 1.54 Å

2. Oxygen gas liquefies at −183.0°C and freezes at −218.4°C. Work out its melting point, T_m, and boiling point, T_b, in kelvin. (Section 1.2)

Strategy

Apply Equation 1.1to convert the temperatures from units of °C to K.

Solution

Applying Equation 1.1,

$$T/K = \theta/°C + 273.15$$

then, for oxygen,

$$T_m = (-218.4°C/°C + 273.15) \text{ K} = 54.8 \text{ K}$$
$$T_b = (-183.0°C/°C + 273.15) \text{ K} = 90.2 \text{ K}$$

3. A sealed flask holds 10 dm^3 of gas. What is this volume in (a) cm^3, (b) m^3, (c) litres? (Section 1.2)

Strategy

Use the SI prefixes in Table 1.3.

Solution

(a) 1 dm^3 = 1000 cm^3

So, 10 dm^3 = 10 000 cm^3 (1×10^4 cm^3)

(b) $10 \text{ dm}^3 = 10 \text{ (dm)}^3 = 10 \times (10^{-1}\text{m})^3 = 10 \times 10^{-3}\text{m}^3 = 1 \times 10^{-2}\text{m}^3$

$$= 0.01 \text{ m}^3$$

(c) $1 \text{ dm}^3 = 1 \text{ L}$, so $10 \text{ dm}^3 = 10 \text{ L}$

4.   What is the SI derived unit for the speed of a molecule? (Section 1.2)

Strategy

Write an expression for speed, and substitute the appropriate SI base units.

Solution

The speed of a molecule may be calculated as

$$\text{speed} = \text{distance travelled/time}$$

so that

$$\text{SI unit of speed} = \text{SI unit of length/SI unit of time} = \text{m/s} = \text{m s}^{-1}$$

5.   How many moles of atoms are contained in the following masses: (a) 22.0 g of magnesium; (b) 43.2 g of chlorine (Cl); (c) 126 mg of gold; (d) 1.00 kg of mercury? (Section 1.3)

Strategy

Use Equation 1.2 to determine the amount of each element, using the SI prefixes in Table 1.3 to ensure that the units are consistent.

$$\text{amount (in mol)} = \frac{\text{mass (in g)}}{\text{molar mass (in g mol}^{-1})} \qquad (1.2)$$

Solution

(a)   Amount of magnesium $= \dfrac{22.0 \text{ g}}{24.31 \text{ g mol}^{-1}} = 0.905$ mol of Mg atoms

(b)   Amount of chlorine $= \dfrac{43.2 \text{ g}}{35.45 \text{ g mol}^{-1}} = 1.22$ mol of Cl atoms

(c)　Amount of gold　$= \dfrac{0.126\ g}{196.97\ g\,mol^{-1}} = 6.40 \times 10^{-4}$ mol of Au atoms

(d)　Amount of mercury　$= \dfrac{1000\ g}{200.59\ g\,mol^{-1}} = 4.99$ mol of Hg atoms

6.　Calculate the amount of each substance contained in the following masses: (a) 89.2 g of carbon dioxide (CO_2); (b) 43.2 g of chlorine (Cl_2); (c) 0.48 kg of calcium hydroxide ($Ca(OH)_2$); (d) 25 tonnes of water, H_2O (1 tonne = 1×10^6 g). (Section 1.3)

Strategy

Use Equation 1.2 to determine the amount of each substance, using the SI prefixes in Table 1.3 to ensure that the units are consistent.

$$\text{amount (in mol)} = \frac{\text{mass (in g)}}{\text{molar mass (in g mol}^{-1})} \tag{1.2}$$

Solution

(a)　$M_r\,(CO_2) = 12.01 + (2 \times 16.00) = 44.01$

　　Amount of $CO_2 = \dfrac{89.2\ g}{44.01\ g\,mol^{-1}} = 2.03$ mol

(b)　$M_r\,(Cl_2) = (2 \times 35.45) = 70.90$

　　Amount of $Cl_2 = \dfrac{43.2\ g}{70.90\ g\,mol^{-1}} = 0.609$ mol

(c)　$M_r\,(Ca(OH)_2) = 40.08 + (2 \times 17.01) = 74.10$

　　Amount of $Ca(OH)_2 = \dfrac{480\ g}{74.10\ g\,mol^{-1}} = 6.48$ mol

(d)　$M_r\,(H_2O) = 16.00 + (2 \times 1.01) = 18.02$

　　Amount of $H_2O = \dfrac{25 \times 10^6\ g}{18.02\ g\,mol^{-1}} = 1.4 \times 10^6$ mol

7.　What is the mass (in g) of (a) 5.46 mol of CuO; (b) 0.107 mol of $KMnO_4$; (c) 2.85 mmol of C_2H_5OH; (d) 1.95 μmol of HCN? (Section 1.3)

Strategy

Use Equation 1.2 to determine the amount of each compound, using the SI prefixes in Table 1.3 to ensure that the units are consistent.

$$\text{amount (in mol)} = \frac{\text{mass (in g)}}{\text{molar mass (in g mol}^{-1})} \qquad (1.2)$$

Solution

(a) $M_r(CuO) = 63.55 + 16.00 = 79.55$

Mass of CuO $= 79.55 \text{ g mol}^{-1} \times 5.46 \text{ mol} = 434 \text{ g}$

(b) $M_r(KMnO_4) = 39.10 + 54.94 + (4 \times 16.00) = 158.04$

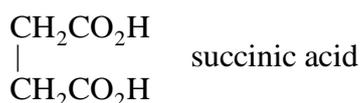Mass of $KMnO_4 = 158.04 \text{ g mol}^{-1} \times 0.107 \text{ mol} = 16.9 \text{ g}$

(c) $M_r(C_2H_5OH) = (2 \times 12.01) + (6 \times 1.01) + 16.00 = 46.08$

Mass of $C_2H_5OH = 46.08 \text{ g mol}^{-1} \times (2.85 \times 10^{-3} \text{ mol}) = 0.131 \text{ g}$

(d) $M_r(HCN) = 1.01 + 12.01 + 14.01 = 27.03$

Mass of HCN $= 27.03 \text{ g mol}^{-1} \times (1.95 \times 10^{-6} \text{ mol}) = 5.27 \times 10^{-5} \text{ g}$

8. The structure of succinic acid is shown below

$$\begin{array}{l} CH_2CO_2H \\ | \\ CH_2CO_2H \end{array} \quad \text{succinic acid}$$

(a) Write down the molecular formula of succinic acid and work out its molar mass.

(b) What is the empirical formula of succinic acid?

(c) What is the percentage of carbon in succinic acid?

(d) Calculate the amount of succinic acid in a 0.125 g sample of the pure acid.

(e) How many molecules of succinic acid are present in the 0.125 g sample?

(f) How many carbon atoms are present in the 0.125 g sample? (Section 1.3)

Strategy

(a) and (b) Find the molecular formula of succinic acid by counting the numbers of each type of atom in the molecule. The empirical formula is the simplest formula that gives the ratio of the numbers of different types of atoms in the molecule.

(c) Find the ratio of the relative mass of carbon in the molecule to the total relative

molecular mass of succinic acid to find the percentage of carbon.

(d)-(f) Use Equation 1.2 to find the amount in moles from the mass of the sample and multiply by the Avogadro constant, N_A, to work out the number of molecules. To find the number of carbon atoms in the sample, multiply the number of succinic acid molecules by the number of carbon atoms in each molecule.

Solution

(a) Molecular formula $C_4H_6O_4$

$M_r(C_4H_6O_4) = (4 \times 12.01) + (6 \times 1.01) + (4 \times 16.00) = 118.1$

Molar mass, $M = 118.1$ g mol^{-1}

(b) Empirical formula $C_2H_3O_2$

(c) $\%C = \dfrac{(4 \times 12.01) \text{ g mol}^{-1}}{118.1 \text{ g mol}^{-1}} \times 100\% = 40.7\%$

(d) Amount of succinic acid $= \dfrac{0.125 \text{ g}}{118.1 \text{ g mol}^{-1}} = 1.06 \times 10^{-3}$ mol

(e) Number of molecules in 0.125 g succinic acid

$= N_A \times (1.06 \times 10^{-3} \text{ mol})$

$= (6.022 \times 10^{23} \text{ mol}^{-1}) \times (1.06 \times 10^{-3} \text{ mol})$

$= 6.38 \times 10^{20}$

(f) Number of C atoms in 0.125 g succinic acid

$= 4 \times (6.38 \times 10^{20})$

$= 2.55 \times 10^{21}$

9. Write down the formulae for sodium azide and sodium nitride. Use these to explain the difference between the N_3^- ion and the N^{3-} ion. (Section 1.4)

Strategy

Think about the difference the position of the number '3' makes to the structure of the ions.

Solution

Sodium azide NaN_3; sodium nitride Na_3N

In the azide ion, three nitrogen atoms are linked by covalent bonds. The whole unit has a single negative charge, $(N_3)^-$, and so combines with one Na^+ ion in sodium azide.

The nitride ion consists of a single nitrogen with a charge of 3−, so it combines with three Na^+ ions in sodium nitride.

10. The fertilizer, ammonium nitrate, is made by reacting ammonia with nitric acid.

(a) Write a balanced equation, with state symbols, for the reaction of ammonia gas with nitric acid to form a solution of ammonium nitrate.

(b) Rewrite the equation to show the ions present in the reactants and products, and hence write an ionic equation for the reaction.

(c) What *type* of chemical reaction does your equation represent?

Nitric acid is manufactured from ammonia. The first stage in this process involves burning ammonia in oxygen on the surface of a platinum gauze catalyst. The products are NO and H_2O.

(d) Construct a balanced equation for the burning of NH_3 in O_2.

(e) Calculate the maximum mass of NO that could be obtained by burning 1.00 kg of NH_3 and the mass of oxygen required.

(f) In practice, 1.45 kg of NO were obtained from the reaction in (e). What was the percentage yield of the reaction? (Section 1.4)

Strategy

(e) Use the stoichiometry of the balanced chemical equation to determine the amounts of NH_3 and O_2 that react exactly, and hence the mass of NO formed and the mass of O_2 required when 1.00 kg of NH_3 is burnt.

(f) Use Equation 1.4 to calculate the percentage yield.

Solution

(a) NH_3 (g) + HNO_3 (aq) → NH_4NO_3 (aq)

(b) Expressing all of the components in terms of their constituent ions
NH_3 (g) + H^+ (aq) + $\cancel{NO_3^-}$ (aq) → NH_4^+ (aq) + $\cancel{NO_3^-}$ (aq)
and then crossing out the spectator ions
NH_3 (g) + H^+ (aq) → NH_4^+ (aq)

(c) The reaction may be classed as an acid–base reaction (see p.23). The HNO_3 (aq) yields H^+ (aq) ions and so is the acid, whilst NH_3 (aq) is the base.

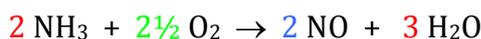(d) First write an unbalanced equation,

$$NH_3 + O_2 → NO + H_2O$$

Balance the number of hydrogen atoms by inserting a 2 before NH_3 and a 3 before H_2O

$$2\ NH_3 + O_2 \rightarrow NO + 3\ H_2O$$

Balance the number of nitrogen atoms by inserting a 2 before NO.

$$2\ NH_3 + O_2 \rightarrow 2\ NO + 3\ H_2O$$
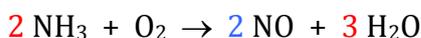
Balance the number of oxygen atoms by inserting a 2½ before O_2.
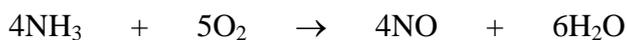
$$2\ NH_3 + 2½\ O_2 \rightarrow 2\ NO + 3\ H_2O$$

The equation is now balanced. You can, if you wish, multiply through by 2 to remove the fraction. The equation then becomes

$$4\ NH_3 + 5\ O_2 \rightarrow 4\ NO + 6\ H_2O$$

(Either of the balanced equations can be used in (e).)

(e)   Using the general approach in Worked Example 1.6

$4NH_3$	$+$	$5O_2$	$\rightarrow$	$4NO$	$+$	$6H_2O$

4 mol   5 mol   4 mol

4×17.04 g 5×32.00 g 4×30.01 g

$= 68.16$ g $= 160.0$ g $= 120.04$ g

1.00 g $\dfrac{160.0}{68.16}$ g $\dfrac{120.04}{68.16}$ g

1.00 kg   2.35 kg   1.76 kg

(f)   From Equation 1.4,

percentage yield $= \dfrac{\text{actual yield}}{\text{theoretical yield}} \times 100\ \%$ (1.4)

$= \dfrac{1.45\ \text{kg}}{1.76\ \text{kg}} \times 100\ \% = 82\%$

11.   Pentanoic acid ($C_4H_9CO_2H$) can be synthesized in three steps from butan-1-ol (C_4H_9OH) as shown below:

C_4H_9OH $\xrightarrow{65\%}$ C_4H_9Cl $\xrightarrow{85\%}$ C_4H_9CN $\xrightarrow{67\%}$ $C_4H_9CO_2H$
butan-1-ol   1-chlorobutane   pentanenitrile   pentanoic acid
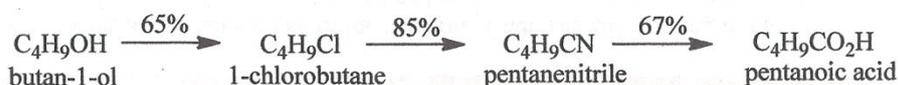
(a) What is the overall yield for the conversion of butan-1-ol to pentanoic acid?

(b) If you carried out the synthesis starting with 20.0 g of butan-1-ol, what mass of pentanoic acid would you obtain?

(c) Even when there are no side reactions in an organic reaction, a 100% yield of product is rarely obtained. Suggest reasons why this might be. (Section 1.4)

Strategy

Use the stoichiometry of the overall reaction to work out the yield of pentanoic acid if the reaction gave 100% yield. Then, use your answer in (a) to work out what mass would actually be obtained.

Solution

(a) $\dfrac{65}{100} \times \dfrac{85}{100} \times \dfrac{67}{100} = \dfrac{370\,175}{1\,000\,000} = \dfrac{37.0}{100}$ which is 37%

(b) $M_r(C_4H_9OH) = (4 \times 12.01) + (10 \times 1.01) + 16.00 = 74.14$

$M_r(C_4H_9CO_2H) = (5 \times 12.01) + (10 \times 1.01) + (2 \times 16.00) = 102.15$

For 100% yield:

1 mol of C_4H_9OH gives 1 mol of $C_4H_9CO_2H$

74.14 g of C_4H_9OH gives 102.15 g of $C_4H_9CO_2H$

20.0 g of C_4H_9OH gives 27.6 g of $C_4H_9CO_2H$

For 37% yield:

20.0 g of C_4H_9OH gives $\dfrac{27.6 \text{ g} \times 37}{100} = 10.2$ g of $C_4H_9CO_2H$

(c) A small amount of product may be lost during isolation from the reaction mixture and during purification (for example, in a solvent extraction, distillation or recrystallization). Also, some reactions do not go to completion, but reach a state of equilibrium when some reactants are still present.

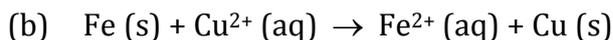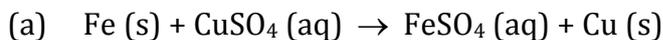**12.** A stream running out from a copper mine contains a dilute solution of copper sulfate. As it passes over an iron grid, copper metal deposits on the grid.

(a) Write a balanced equation, with state symbols, for the reaction taking place.

(b) Write an ionic equation for the reaction.

(c) Assign oxidation states to the elements in each of the reactants and products in

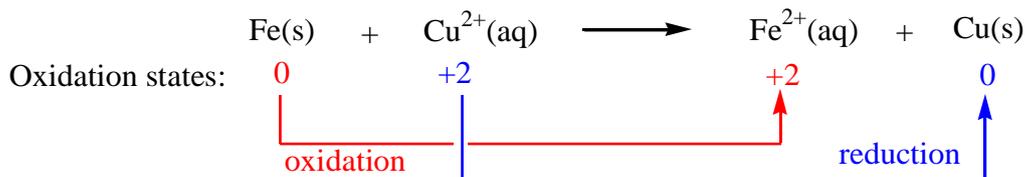the equation in (b). Use these values to decide what has been oxidized and what reduced. (Section 1.4)

Strategy

Follow the approach used in Worked Example 1.10.

Solution

(a)   Fe (s) + $CuSO_4$ (aq) $\rightarrow$ $FeSO_4$ (aq) + Cu (s)

(b)   Fe (s) + Cu^{2+} (aq) $\rightarrow$ Fe^{2+} (aq) + Cu (s)

(c)

$$\text{Fe(s)} \quad + \quad \text{Cu}^{2+}\text{(aq)} \quad \longrightarrow \quad \text{Fe}^{2+}\text{(aq)} \quad + \quad \text{Cu(s)}$$

Oxidation states: 0   +2   +2   0

oxidation   reduction

The oxidation state of iron changes from 0 to +2, as the Fe (s) reacts to form Fe^{2+} (aq). The oxidation state of copper changes from +2 to 0 as the Cu^{2+} (aq) reacts to form Cu(s). The iron is thus oxidised and the copper reduced.

13.   (a)   What are the systematic names for (i) CS_2; (ii) Cl_2O_7; (iii) XeF_6; (iv) $(NH_4)_2SO_4$;

(v) $CrCl_3$; (vi) KIO_4?

(b)   Write the formula of each of the following compounds: (i) sodium sulfite; (ii)

barium carbonate; (iii) iron(II) chloride; (iv) sodium thiosulfate; (v) diiodine

pentoxide; (vi) dinitrogen oxide. (Section 1.4)

Strategy

Use the rules outlined on pp.29-30 and in Worked Example 1.9. Table 1.6 on p.22 gives a list of common ions.

Solution

(a)   (i) Carbon disulfide (sulfur is slightly more electronegative than carbon and its name is changed to end in –ide).
(ii) Dichlorine heptoxide (use prefixes to indicate the numbers of atoms in the formula)
(iii) Xenon hexafluoride
(iv) Ammonium sulphate (the ammonium ion, NH_4^+, has a single positive

charge, so two ammonium ions combine with one sulphate ion, SO_4^{2-}.)

(v) Chromium(III) chloride (Roman numerals in brackets show the oxidation state of chromium)

(vi) potassium periodate (or potassium iodate(VII) (see Table 1.6 p.22 for analogous perchlorate ion ClO_4^-).

(b)   (i) Na_2SO_3; (ii) $BaCO_3$; (iii) $FeCl_2$; (iv) $Na_2S_2O_3$; (v) I_2O_5; (vii) N_2O.

14.   In most compounds, H has an oxidation state of +1 and O has an oxidation state of −2. The following compounds are exceptions to this rule. Assume each metal has the oxidation state of its most common ion and that F has an oxidation state of −1. Find the oxidation state of H and O in each compound: (a) KO_2; (b) Na_2O_2; (c) MgH_2; (d) $LiAlH_4$; (e) OF_2. (Section 1.4)

<u>Strategy</u>

Follow the rules for assigning oxidation states on p.28. All of the compounds are uncharged, so the sum of the oxidation states of the various elements is zero.

<u>Solution</u>

(a) The oxidation state of K is +1, since potassium is most commonly found as K^+ ions. In order that the overall sum of oxidation states is 0, the oxidation state of O must be −½.

(b) In the same way, since sodium is usually found as Na^+, the oxidation state of Na is +1, so the oxidation state of the O must be −1 so that the oxidation states sum to 0.

(c) Magnesium exists as Mg^{2+} ions, so the oxidation state of Mg is +2. The oxidation state of H must therefore be −1.

(d) Lithium exists as Li^+ and aluminium as Al^{3+}. The oxidation states of the Li and Al are therefore +1 and +3, repectively, meaning that the oxidation state of H is −1.

(e) Fluorine forms F^- ions, so that the two F atoms each have an oxidation number of −1. The oxidation number of O must therefore be +2.

15*.   Sodium chromate (Na_2CrO_4) can be prepared by oxidizing a chromium(III) salt with sodium peroxide (Na_2O_2) in alkaline solution. The Cr^{3+} ions are oxidized to CrO_4^{2-} ions. The O_2^{2-} ions are reduced to OH^- ions. Construct half equations and a balanced overall equation for the reaction. (Section 1.4)
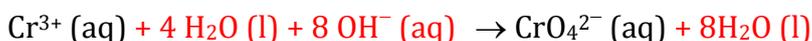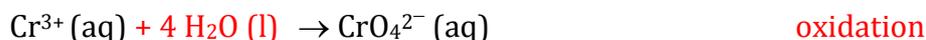
Strategy

Follow the steps for constructing and balancing a redox equation from half equations on p.32.
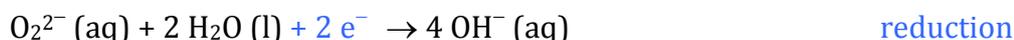
Solution

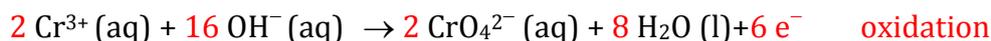**Steps 1and 2**: Write unbalanced half equations, with elements other than O and H balanced

$$Cr^{3+} (aq) \rightarrow CrO_4{}^{2-} (aq) \qquad \text{oxidation}$$

$$O_2{}^{2-} (aq) \rightarrow OH^- (aq) \qquad \text{reduction}$$

Step 3: In alkaline solution, balance O and H by first adding H_2O to balance the O, then balance the H by adding H_2O to the side of the half equation that needs H, whilst adding OH^- to the other side. Cancelling out any surplus water.
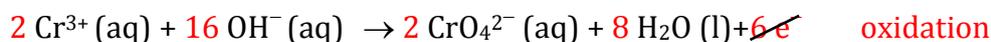
$$Cr^{3+} (aq) + 4\ H_2O\ (l) \rightarrow CrO_4{}^{2-} (aq) \qquad \text{oxidation}$$

$$Cr^{3+} (aq) + 4\ H_2O\ (l) + 8\ OH^-\ (aq) \rightarrow CrO_4{}^{2-} (aq) + 8H_2O\ (l)$$

$$Cr^{3+} (aq) + 8OH^-\ (aq) \rightarrow CrO_4{}^{2-} (aq) + 4H_2O\ (l)$$

$$O_2{}^{2-} (aq) \rightarrow 2\ OH^-\ (aq) \qquad \text{reduction}$$

$$O_2{}^{2-} (aq) + 2\ H_2O(l) \rightarrow 4\ OH^-\ (aq)$$

Step 4: Balance electric charges by adding electrons

$$Cr^{3+} (aq) + 8\ OH^-\ (aq) \rightarrow CrO_4{}^{2-} (aq) + 4\ H_2O\ (l) + 3\ e^- \qquad \text{oxidation}$$

$$O_2{}^{2-} (aq) + 2\ H_2O\ (l) + 2\ e^- \rightarrow 4\ OH^-\ (aq) \qquad \text{reduction}$$
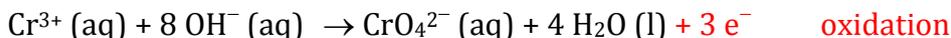
Step 5: Multiplying the oxidation equation ×2, and the reduction equation ×3, so that the number of electrons transferred is the same in both half equations.

$$2\ Cr^{3+} (aq) + 16\ OH^-\ (aq) \rightarrow 2\ CrO_4{}^{2-} (aq) + 8\ H_2O\ (l) + 6\ e^- \qquad \text{oxidation}$$

$$3\ O_2{}^{2-} (aq) + 6\ H_2O\ (l) + 6\ e^- \rightarrow 12\ OH^-\ (aq) \qquad \text{reduction}$$

Step 6: Add the two half equations and cancel out the electrons

$$2\ Cr^{3+} (aq) + 16\ OH^-\ (aq) \rightarrow 2\ CrO_4{}^{2-} (aq) + 8\ H_2O\ (l) + \cancel{6\ e^-} \qquad \text{oxidation}$$

$$3\ O_2{}^{2-} (aq) + 6\ H_2O\ (l) + \cancel{6\ e^-} \rightarrow 12\ OH^-\ (aq) \qquad \text{reduction}$$

$$2Cr^{3+}(aq) + 16OH^-(aq) + 3O_2{}^{2-}(aq) + 6H_2O(l) \rightarrow 2CrO_4{}^{2-}(aq) + 8H_2O(l) + 12OH^-(aq) \quad \text{overall}$$

Steps 7 and 8: Simplify the equation by cancelling out $OH^-(aq)$ ions and $H_2O(l)$ molecules. The atoms and charges on each side of the equation balance.

$$2\ Cr^{3+}(aq) + 4\ OH^-\ (aq) + 3\ O_2{}^{2-}(aq) \rightarrow 2\ CrO_4{}^{2-}\ (aq) + 2\ H_2O(l) \qquad \text{overall}$$

16*. To prepare a very dilute solution, it is more accurate to make up a more concentrated standard solution, and carry out a series of successive dilutions, than to weigh out a very small mass of the solute.

A solution was made by dissolving 0.587 g of $KMnO_4$ in dilute sulfuric acid and making the volume of solution up to 1 dm^3 in a volumetric flask. 10.0 cm^3 of this solution were transferred to a second 1 dm^3 volumetric flask and diluted to the mark with water. The dilution process was then repeated, i.e. 10.0 cm^3 of this solution were transferred to a 1 dm^3 volumetric flask and diluted to the mark with water.

(a) What mass (in mg) of $KMnO_4$ would you have had to weigh out to make 500 cm^3 of a solution with the same concentration as the final dilute solution?

(b) What is the concentration of the final dilute $KMnO_4$ solution in mol dm^{-3}?

(Section 1.5)

Strategy

(a) Calculate the dilution factor (how much the solution is diluted each time). From the mass of $KMnO_4$ in 1000 cm^3 of the initial solution, work out the mass of $KMnO_4$ in 1000 cm^3 of the final solution, and hence the mass of $KMnO_4$ in 500 cm^3 of the final solution.

(b) Work out the molar mass of $KMnO_4$ and then use Equation 1.2 on p.19 (or Equation 1.8 on p.35) to calculate the molar concentration.

Solution

(a) In each dilution, 10 cm^3 is taken from 1000 cm^3 in the volumetric flask and made up to 1000 cm^3, so each dilution is

$$\frac{10 \text{ cm}^3}{1000 \text{ cm}^3} = \frac{1}{100} = 0.01$$

The dilution is carried out twice, so is equivalent to

$$\frac{1}{100} \times \frac{1}{100} = 0.0001 = 1 \times 10^{-4}$$

Mass of $KMnO_4$ in final solution $= \dfrac{0.587 \text{ g}}{100 \times 100}$

$$= 5.87 \times 10^{-5} \text{ g}$$

Mass of $KMnO_4$ needed to make 500 cm^3 of this solution

$$= \frac{5.87 \times 10^{-5} \text{ g}}{2} = 2.94 \times 10^{-5} \text{ g}$$

$$= 0.0294 \text{ mg}$$

(b) $M_r(KMnO_4) = 39.10 + 54.94 + (4 \times 16.00) = 158.04$

Using Equation 1.2 on p. 19,

amount of $KMnO_4$ in 1000 cm^3 final solution $= \dfrac{\text{mass (in g)}}{\text{molar mass (in g mol}^{-1})}$

$$= \frac{5.87 \times 10^{-5} \, \cancel{g}}{158.04 \, \cancel{g} \, \text{mol}^{-1}}$$

$$= 3.71 \times 10^{-7} \text{ mol}$$

Concentration of final solution $= 3.71 \times 10^{-7} \text{ mol dm}^{-3}$

17* The most common type of kidney stone is formed from calcium ethanedioate (CaC_2O_4) which precipitates out in the urinary tract when concentrations of Ca^{2+} ions and $C_2O_4{}^{2-}$ ions become too high. Magnesium ions are known to inhibit the formation of kidney stones. To analyse the concentrations of Ca^{2+} (aq) and Mg^{2+} (aq) in a sample of urine, the ions were precipitated as ethandioates and the mixed precipitate of CaC_2O_4 and MgC_2O_4 analysed by gravimetric analysis.

The solid ethanedioates were first heated to form a mixture of $CaCO_3$ and MgO. The mass of this mixture was 0.0433 g. This solid was then heated more strongly to give a mixture of CaO and MgO. The mass of the solid residue from this process was 0.0285 g. What was the mass of the Ca^{2+} ions in the original sample of urine? (Section 1.5)

Strategy

Calculate the mass, and therefore the amount, of CO_2 released when the $CaCO_3$ is heated strongly. Use the stoichiometry of the decomposition reaction of $CaCO_3$ to determine the amount of $CaCO_3$ in the mixture of $CaCO_3$ and MgO. You can now use Equation 1.2 on p.19 to find the mass of Ca^{2+} ions present in the original sample of urine.

Solution

Mass of $CaCO_3 + MgO = 0.0433$ g

On heating strongly,

$CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$

1 mol   1 mol

Mass of CO_2 lost on heating strongly $= 0.0433$ g $- 0.0285$ g $= 0.0148$ g

$M_r(CO_2) = 12.01 + 32.00 = 44.01$

Using Equation 1.2 on p.19,

$$\text{amount of } CO_2 \text{ lost} = \frac{\text{mass (in g)}}{\text{molar mass (in g mol}^{-1})}$$

$$= \frac{0.0148\,\cancel{g}}{44.01\,\cancel{g}\,\text{mol}^{-1}}$$

$$= 3.363 \times 10^{-4}\,\text{mol}$$

Since 1 mol of CO_2 is liberated from 1 mol $CaCO_3$, the amount of $CaCO_3$ in the mixture of $CaCO_3$ and MgO

$$= 3.363 \times 10^{-4}\,\text{mol}$$

$A_r(Ca) = 40.08$

Using Equation 1.2,

$$\text{mass of } Ca^{2+} \text{ ions} = (3.363 \times 10^{-4}\,\cancel{\text{mol}}) \times (40.08\,\text{g}\,\cancel{\text{mol}^{-1}})$$

$$= 0.0135\,\text{g}$$

This was the mass of Ca^{2+} ions in the original sample of urine.

18*. The standard enthalpy change of combustion of heptane, C_7H_{16}, at 298 K, is −4817 kJ mol^{-1}.

(a) Write a thermochemical equation for the complete combustion of heptane to carbon dioxide and water.

(b) What is the enthalpy change when 50 g of heptane are burned?

(c) What mass of heptane would be needed to provide 100 MJ of energy? (Section 1.6)

Strategy

(a) Write a balanced equation for the complete combustion of 1 mol of heptane and give the enthalpy change corresponding to this process.

(b) Work out the molar mass of heptane. From the enthalpy change when this mass of heptane is burned, you can calculate the enthalpy change when 50 g are burned.

(c) You know the mass of heptane needed to liberate 4817×10^3 J of energy, so you can calculate the mass needed to liberate 100MJ of energy. (1 MJ = 1×10^6 J)

Solution

(a) $C_7H_{16}(l) + 11O_2(g) \rightarrow 7CO_2(g) + 8H_2O(l)$ $\Delta_cH^{\oplus} = -4817$ kJ mol^{-1}
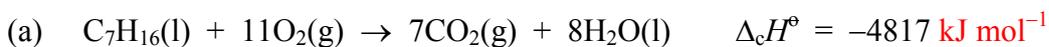
(b) $M_r(C_7H_{16}) = 100$

Enthalpy change when 100 g of C_7H_{16} are burned is –4817 kJ

So, enthalpy change when 50 g of C_7H_{16} are burned is –2409 kJ

(c) 4817×10^3 J of energy are liberated from 100 g of heptane

1 J of energy is liberated from $\left(\dfrac{100}{4817 \times 10^3}\right)$ g of heptane

100×10^6 J of energy from $\left(\dfrac{100 \times (100 \times 10^6)}{4817 \times 10^3}\right)$ g

$= 2.08 \times 10^3$ g $= 2.08$ kg of heptane

19*. (a) List the non-covalent interactions present in liquid water. Which is responsible for the strongest interactions between the molecules?

(b) Explain why the value of $\Delta_{vap}H^{\oplus}(H_2O)$ is unusually high for a molecule of its size.

(c) In a storm, 3 cm of rain falls on the city of Leeds, which has an area of approximately 500 km². Estimate the energy released as heat when this quantity of water condenses from vapour to form rain. (Density of water is 1.00 g cm^{-3}; $\Delta_{vap}H^{\oplus}(H_2O) = +40.7$ kJ mol^{-1} at 298 K.)

(d) The output from a large 2000 MW power station is 2000 MJ s^{-1}. How long would it take the power station to deliver the same quantity of energy as was released by the condensation of the rain in (c)? (Sections 1.7 and 1.8)

Strategy

(c) Calculate the volume of water that falls on the city and, using the density of water, find the mass of this volume of water. Use Equation 1.2 on p.19 to determine the amount in mol of the water and hence the energy released when it condenses.

Solution

(a) Within liquid water, there are London dispersion interactions, dipole-dipole interactions and hydrogen bonding. Hydrogen bonding is responsible for the strongest attraction between the molecules.

(b) Energy must be supplied to overcome intermolecular attractions due to extensive hydrogen bonding in liquid water.

(c) $1 \text{ km} = 1 \times 10^5 \text{ cm}$, so area of Leeds is $500 \times 10^{10} \text{ cm}^2$

Volume of water $= (500 \times 10^{10} \text{ cm}^2) \times 3 \text{ cm}$

$= 1.5 \times 10^{13} \text{ cm}^3$

Volume of a cylinder $=$ area of the cross section $\times$ length

Density of water $= 1.00 \text{ g cm}^{-3} =$ mass/volume

So, mass of water $= 1.00 \text{ g cm}^{-3} \times (1.5 \times 10^{13} \text{ cm}^3) = 1.5 \times 10^{13} \text{ g}$

$M_r(\text{H}_2\text{O}) = 18.0$

Using Equation 1.2 on p.19,

$$\text{Amount of water} = \frac{\text{mass (in g)}}{\text{molar mass (in g mol}^{-1})} \qquad (1.2)$$

$$= \frac{1.5 \times 10^{13} \text{ g}}{18.0 \text{ g mol}^{-1}}$$

$$= 8.33 \times 10^{11} \text{ mol}$$

Heat energy released $= (8.33 \times 10^{11} \text{ mol}) \times (40.7 \text{ kJ mol}^{-1})$

$= 3.39 \times 10^{13} \text{ kJ}$

$= 3 \times 10^{16} \text{ J}$

(d) Time for the power station to deliver this amount of energy

$$= \frac{3 \times 10^{16} \text{ J}}{2000 \times 10^6 \text{ J s}^{-1}}$$

$$= 1.5 \times 10^7 \text{ s}$$

$$= 174 \text{ days (about 6 months)}$$

20[*]. Nitrogen dioxide gas is heated in a sealed container at 700 K until the system comes to equilibrium. The nitrogen dioxide dissociates into nitrogen monoxide and oxygen in an endothermic process:

$$2\text{NO}_2(g) \rightleftharpoons 2\text{NO}(g) + \text{O}_2(g)$$

The equilibrium constant at 700 K is $2.78 \times 10^{-2} \text{ mol dm}^{-3}$.

(a) Write an expression for K_c.

(b) State how the position of equilibrium would be affected by:

(i) an increase in temperature   (ii) an increase in the total pressure.

(c) At equilibrium at 700 K, the concentration of nitrogen monoxide was found to be 0.017 $mol\ dm^{-3}$. What was the concentration of nitrogen dioxide in the equilibrium mixture? (Section 1.9)

Strategy

(a) and (b) Use Equation 1.23 on p.60 to write an expression for K_c and Le Chatelier's Principle on p.57 to decide how the position of equilibrium is affected by an increase in temperature and by an increase in the total pressure.

(c) Use the stoichiometric equation to work out the concentration of $O_2(g)$ at equilibrium, and then substitute values into the expression for K_c.

Solution

(a) $K_c = \dfrac{[NO]^2[O_2]}{[NO_2]^2}$

(b) (i) Forward reaction is endothermic, so, if the temperature increases, the position of equilibrium moves to the right.

(ii) There are 2 molecules of gaseous reactants on the left hand side of the equation, and 3 molecules of gaseous products on the right hand side. An increase in the total pressure causes the position of equilibrium to move to the left.

(c) At the start, the reaction vessel contained pure NO_2.

From the equation, at equilibrium: $[O_2] = 0.5 \times [NO] = 0.0085\ mol\ dm^{-3}$

$$K_c = 2.78 \times 10^{-2}\ mol\ dm^{-3} = \frac{\left(0.017\ mol\ dm^{-3}\right)^2 \times \left(0.0085\ mol\ dm^{-3}\right)}{[NO_2]^2}$$

$$[NO_2]^2 = \frac{\left(0.017\ mol\ dm^{-3}\right)^2 \times \left(0.0085\ \cancel{mol\ dm^{-3}}\right)}{2.78\times10^{-2}\ \cancel{mol\ dm^{-3}}}$$

$[NO_2] = 0.0094\ mol\ dm^{-3}$

21*. One stage in the manufacture of methanol from methane involves the conversion of synthesis gas (a mixture of CO and H_2) to methanol. The conversion is carried out over a catalyst at a temperature of around 500 K and a pressure of 100 atm.

$$CO\ (g) + 2H_2\ (g) \ \rightleftharpoons\ CH_3OH\ (g) \qquad \Delta H = -90.7\ kJ\ mol^{-1}$$
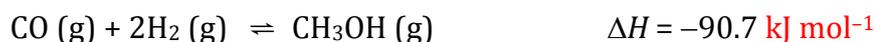
(a) Write an expression for K_p for the reaction.

(b) At 500 K and 100 atm pressure, an equilibrium mixture contains 42% CH_3OH and 48% CO. Calculate a value for K_p at this temperature.

(c) Use Le Chatelier's principle to predict what would happen to the percentage of methanol in the mixture if: (i) the temperature increases; (ii) the pressure increases; (iii) hydrogen is added at constant temperature and pressure. (Section 1.9)

Strategy

(a)   Use Equation 1.24 on p.62 to write an expression for K_p.

(b)   Calculate the partial pressures of each of the components from their proportions in the mixture and the total pressure following the method on p.63. Remember that H_2 makes up the remainder of the mixture.

Solution

(a) $K_p = \dfrac{p_{CH_3OH}}{p_{CO} \, p_{H_2}{}^2}$

(b)   Total pressure $= 100 \text{ atm} = p_{CH_3OH} + p_{CO} + p_{H_2}$

Percentage of H_2 in equilibrium mixture must be 10%

$p_{CH_3OH} = 0.42 \times 100 \text{ atm} = 42 \text{ atm}$

$p_{CO} = 0.48 \times 100 \text{ atm} = 48 \text{ atm}$

$p_{H_2} = 0.10 \times 100 \text{ atm} = 10 \text{ atm}$

$K_p = \dfrac{42 \text{ atm}}{(48 \text{ atm}) \times (10 \text{ atm})^2}$

$= 8.8 \times 10^{-3} \text{ atm}^{-2}$

(c)   (i)   Forward reaction is exothermic, so, if the temperature increases, the position of equilibrium moves to the left. The percentage of methanol in the mixture decreases.

(ii)   There are 3 molecules of gaseous reactants on the left hand side of the equation, but only 1 molecule of gaseous product on the right hand side. An increase in pressure causes the position of equilibrium to move to the right. The percentage of methanol in the mixture increases.

(iii)   An increase in the partial pressure of hydrogen will cause the position of equilibrium to move to the right. The percentage of methanol in the mixture increases.