Edexcel International AS Chemistry

Your notes

Redox Chemistry & Acid-Base Titrations

Contents

Oxidation Numbers – Introduction

- There are three definitions of **oxidation** and **reduction** used in different branches of chemistry
- **Oxidation** and **reduction** can be used to describe any of the following processes

Definitions and Examples of Oxidation & Reduction

Oxidation	Reduction
Addition of oxygen e.g. $2Mg + O_2 \longrightarrow 2MgO$	Loss of oxygen e.g. $2CuO + C \longrightarrow 2Cu + CO_2$
Loss of hydrogen e.g. $CH_3OH \xrightarrow{[O]} CH_2O + H_2O$	Addition of hydrogen e.g. $C_2H_4 + H_2 \longrightarrow C_2H_6$
Loss of electrons e.g. $Al \longrightarrow Al^{3+} + 3e^-$	Gain of electrons e.g. $F_2 + 2e^- \longrightarrow 2F^-$

O	OXIDATION	R	REDUCTION
I	IS	I	IS
L	LOSS (OF ELECTRONS)	G	GAIN (OF ELECTRONS)

Use the acronym "Oil Rig" to help you remember the definitions of oxidation and reduction

Oxidation Number

- The oxidation number of an atom is the charge that would exist on an individual atom if the bonding were completely ionic
- It is like the electronic 'status' of an element
- Oxidation numbers are used to
    - Tell if oxidation or reduction has taken place
    - Work out what has been oxidised and/or reduced

- Construct half equations and balance redox equations

Oxidation Numbers of Simple Ions

Atoms	Na in Na = 0	neutral already, no need to add any electrons
Cations	Na in Na^+ = +1	need to add 1 electron to make Na+ neutral
Anions	Cl in Cl^- = −1	need to take 1 electron away to make Cl^- neutral

Worked Example

What are the oxidation numbers of the elements in the following species?

a) C b) Fe^{3+} c) Fe^{2+}

d) O^{2-} e) He f) Al^{3+}

Answers:

a) 0 b) +3 c) +2

d) -2 e) 0 f) +3

- So, in simple ions, the oxidation numbers of the atom is the charge on the ion:
  - Na^+, K^+, H^+ all have an oxidation number of +1
  - Mg^{2+}, Ca^{2+}, Pb^{2+} all have an oxidation number of +2
  - Cl^-, Br^-, I^- all have an oxidation number of -1
  - O^{2-}, S^{2-} all have an oxidation number of -2

Oxidation Number Rules

- A few simple rules help guide you through the process of determining the oxidation number of any element
- Remember, you are determining the oxidation number of a *single* atom
- The oxidation number (ox.no.) refers to a *single* atom in a compound

Oxidation Number Rules Table

Rule	Example
1. The ox.no. of any uncombined element is zero	H_2 Zn O_2
2. Many atoms or ions have fixed ox.no. in compounds	Group 1 elements are always +1 Group 2 elements are always +2 Fluorine is always −1 Hydrogen is +1 (except for in metal hydrides like NaH, where it is −1) Oxygen is −2 (except in peroxides, where it is −1 and in F_2O where it is +2)
3. The ox.no. of an element in a mono−atomic ion is always the same as the charge	Zn^{2+} ox.no. = +2 Fe^{3+} ox.no. = +3 Cl^{-} ox.no. = −1
4. The sum of the ox.no. in a compound is zero	NaCl ox.no. of Na = +1 ox.no. of Cl = −1 Sum ox.no. = 0
5. The sum of ox.no. in an ion is equal to the charge on the ion	SO_4^{2-} ox.no. of S = +6 ox.no. of four O atoms = 4 x (−2) sum ox.no. = −2
6. In either a compound or an ion, the more electronegative element is given the ox.no.	F_2O ox.no. of both F atoms = 2 x (−1) ox.no. of O = +2

Your notes

Roman Numerals

- **Transition metals** are characterised by having **variable oxidation numbers**.
- **Oxidation numbers** can be used in the names of compounds to indicate which **oxidation number** a particular element in the compound is in
- Where the element has a **variable oxidation number**, the number is written afterwards in **Roman numerals**.
- This is called the **STOCK NOTATION** (after the German inorganic chemist Alfred Stock), but is not as widely used for non-metals, so SO_2 is often referred to as sulfur dioxide rather than sulfur(IV) oxide
- For example, iron can be both +2 and +3 so **Roman numerals** are used to distinguish between them

- Fe^{2+} in FeO can be written as **iron(II) oxide**
- Fe^{3+} in Fe_2O_3 can be written as **iron(III) oxide**

Your notes

- More complicated examples include other atoms / ions as part of the formula
  - Potassium manganate(VII) implies that the manganese is in a +7 oxidation
  - Potassium manganate(VII) contains the potassium ion K^+ and the manganate ion MnO_4^-
    - Since the oxygen in the manganate ion is in the -2 oxidation state, there is a total of -8 from the oxygen
    - The manganate ion has an overall -1 charge, which means that the manganese ion must be in the +7 oxidation state

Oxidation Numbers – Calculations

Molecules or Compounds

- In molecules or compounds, the sum of the oxidation numbers on the atoms is zero

Oxidation Number in Molecules or Compounds Table

Elements	H in $H_2 = 0$	Both are the same and must add up to zero
Compounds	C in $CO_2 = +4$	$1 \times (+4)$ and $2 \times (-2) = 0$
	O in $CO_2 = -2$	

- Because CO_2 is a neutral molecule, the sum of the oxidation states must be zero
- For this, one element must have a positive oxidation number and the other must be negative

How do you determine which is the positive one?

- The more electronegative species will have the negative value
- Electronegativity increases across a period and decreases down a group
- O is further to the right than C in the periodic table so it has the negative value

How do you determine the value of an element's oxidation number?

- From its position in the periodic table and / or the other element(s) present in the formula

Hydrogen in metal hydrides, H^-

- An example of a metal hydride is sodium hydrogen, NaH which is a neutral compound

- The hydrogen atom is present as a **hydride ion**
  - This has the symbol H^- and therefore the oxidation state -1
- We can also think of this as a sum
  - Since **group 1** metals always have an oxidation state of **+1** in their compounds, it follows that the hydrogen must have an oxidation state of -1 (+1 -1 = 0)

Oxygen in peroxides, O-

- Peroxides include hydrogen peroxide, H_2O_2, which again is a neutral compound
- The sum of the oxidation states of the hydrogen and oxygen must be **0**
- Since each hydrogen has an oxidation state of +1, each **oxygen must have an oxidation state of –1** to accommodate the neutral charge of the compound

Ions

- The oxidation number of elements within an ion can be calculated using the rules shown in the **Oxidation Number Rules Table** (above)

Worked Example

Deduce the oxidation number of the highlighted element in:

1. The nitrite ion: $\underline{\mathbf{N}}O_2^-$
2. The nitrate ion: $\underline{\mathbf{N}}O_3^-$
3. The thiosulfate ion: $\underline{\mathbf{S}}_2O_3^{2-}$
4. The tetrathionate ion: $\underline{\mathbf{S}}_4O_6^{2-}$

Answers

- In all of the answers, the O is more electronegative than the other element, so the oxidation number of O will be –2

Answer 1:

- The nitrite ion contains two oxygen atoms
  - 2 x –2 = –4
  - The overall ion has an oxidation state of –1, which means that the oxidation numbers of the elements in the ion must add up to –1
  - Therefore, the oxidation number of N is +3
    - N + (2 x –2) = –1 N = +3

Answer 2:

- The nitrate ion contains three oxygen atoms
  - 3 x –2 = –6
  - The overall ion has an oxidation state of –1, which means that the oxidation numbers of the elements in the ion must add up to –1
  - Therefore, the oxidation number of N is +5

Your notes

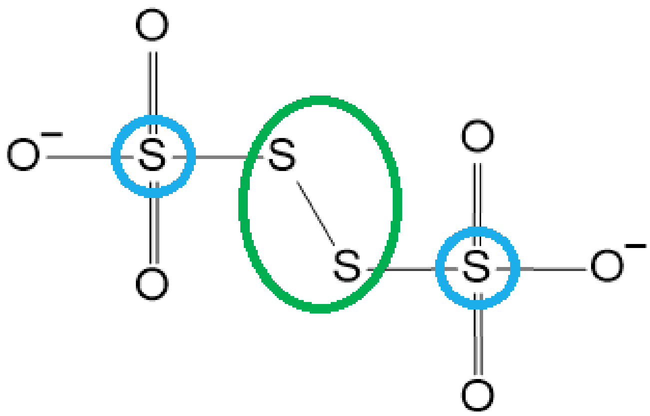
- $N + (3 \times -2) = -1$ $N = +5$

Answer 3:

- The thiosulfate ion contains three oxygen atoms
  - $3 \times -2 = -6$
  - The overall ion has an oxidation state of –2, which means that the oxidation numbers of the elements in the ion must add up to –2
  - The oxidation number of both S atoms is +4
    - $2S + (3 \times -2) = -2$ $2S = +4$
  - Therefore, the oxidation number of S is +2
    - $2S = +4$
    - $S = +2$

Answer 4:

- The tetrathionate ion contains six oxygen atoms
  - $6 \times -2 = -12$
  - The overall ion has an oxidation state of –2, which means that the oxidation numbers of the elements in the ion must add up to –2
  - The oxidation number of all four S atoms is +10
    - $4S + (6 \times -2) = -2$ $4S = +10$
  - Therefore, the oxidation number of S is +2.5
    - $4S = +10$
    - $S = 2.5$
  - This specific question is a rare example showing a fractional oxidation number for the element in question / sulfur
    - This is because the sulfur atoms are in two different environments



    - The fractional oxidation number shows the average oxidation number of the sulfur environments

Your notes

Your notes

Electron Transfer & Oxidation Numbers

- Oxidation and reduction in a reaction can be demonstrated in terms of electron transfer
  - For example:

$$2NH_3 + 3Br_2 \rightarrow N_2 + 6HBr$$

- The oxidation number of N in NH_3 has changed from -3 to 0
  - As the oxidation number has increased, nitrogen has been oxidised
- The oxidation number of Br has changed from 0 to -1
  - As the oxidation number has decreased, bromine has been reduced
- Overall, nitrogen has reduced bromine by donating electrons

Reducing & Oxidising Agents

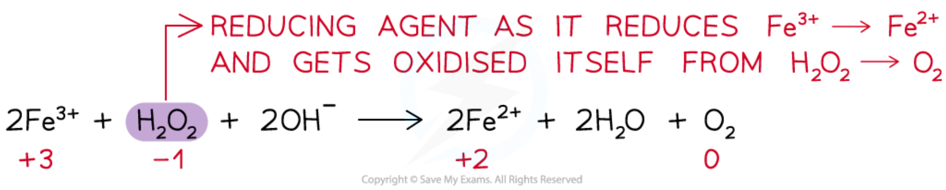
Oxidising agent

- **An oxidising agent** is a substance that **oxidises** another atom or ion by causing it to lose electrons
- An oxidising agent itself gets **reduced – gains electrons**
- Therefore, the **ox. no.** of the oxidising agent **decreases**

Example of an oxidising agent in a chemical reaction

Reducing agent

- **A reducing agent** is a substance that **reduces** another atom or ion by causing it to gain electrons
- A reducing agent itself gets **oxidised – loses / donates electrons**
- Therefore, the **ox. no.** of the reducing agent **increases**



Example of a reducing agent in a chemical reaction

- For a reaction to be recognised as a redox reaction, there must be both an oxidising and reducing agent
- Some substances can act both as oxidising and reducing agents
- Their nature is dependent upon what they are reacting with and the reaction conditions

Worked Example

Four reactions are shown.

In which reaction is the species in bold acting as an oxidising agent?

1. $Cr_2O_7^{2-} + 8H^+ + \mathbf{3SO_3^{2-}} \rightarrow 2Cr^{3+} + 4H_2O + 3SO_4^{2-}$
2. $Mg + \mathbf{Fe^{2+}} \rightarrow Mg^{2+} + Fe$
3. $Cl_2 + \mathbf{2Br^-} \rightarrow 2Cl^- + Br_2$
4. $Fe_2O_3 + \mathbf{3CO} \rightarrow 2Fe + 3CO_2$

Answer:

- The correct option is 2
  - Oxidising agents are substances that oxidise other species, gain electrons and are themselves reduced.
  - Write down the oxidation numbers of each species in the reaction

0		**+2**		**+2**		**0**
Mg	+	Fe^{2+}	→	Mg^{2+}	+	Fe^{-}

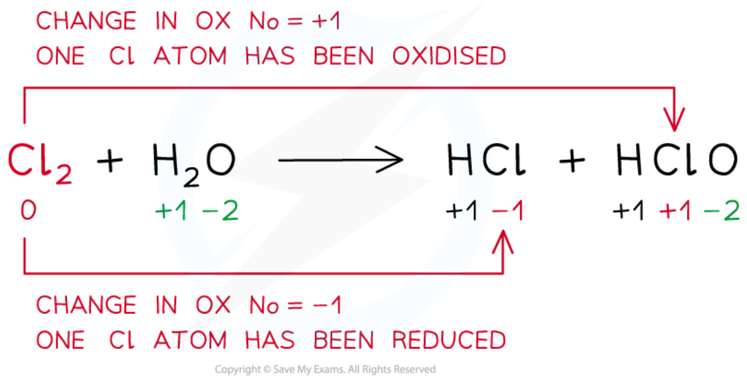
- In equation 2, Fe^{2+} oxidises Mg(0) to Mg^{2+}(+2) and is itself reduced from Fe^{2+}(+2) to Fe(0)

Your notes

Disproportionation

Disproportionation reactions

- A **disproportionation reaction** is a reaction in which the same species is simultaneously oxidised and reduced



Example of a disproportion reaction in which the same species (chlorine in this case) has been both oxidised and reduced

Worked Example

Balancing disproportionation reactions

Balance the disproportionation reaction which takes place when chlorine is added to hot concentrated aqueous sodium hydroxideThe products are Cl^- and ClO_3^- ions and water

Answer

Step 1: Write the unbalanced equation and identify the atoms that change in oxidation number:

Step 2: Deduce the oxidation number changes:

CHANGE IN OX. NO = −1

$$\underset{0}{Cl_2} + \underset{-2\,+1}{OH^-} \longrightarrow \underset{-1}{Cl^-} + \underset{+5\,-2}{ClO_3^-} + \underset{+1\,-2}{H_2O}$$

CHANGE IN OX. NO = +5

Step 3: Balance the oxidation number changes:

CHANGE IN OX. NO = 5 × (−1) = −5

$$\underset{0}{Cl_2} + OH^- \longrightarrow \underset{-1}{5Cl^-} + \underset{+5}{ClO_3^-} + H_2O$$

CHANGE IN OX. NO = 1 × (+5) = +5

Step 4: Balance the charges

$$Cl_2 + \underline{OH^-} \longrightarrow \underline{5Cl^-} + \underline{ClO_3^-} + H_2O$$

- TOTAL ⊕ CHARGE = 0
- TOTAL ⊖ CHARGE = 1−

TOTAL CHARGE REACTANTS = 1−

- TOTAL ⊕ CHARGE = 0
- TOTAL ⊖ CHARGE = 5 × (1−) + (1−) = 6−

TOTAL CHARGE PRODUCTS = 6−

THUS 6 OH^- IONS ARE NEEDED TO BALANCE THE CHARGES ON BOTH SIDES

Step 5: Balance the atoms

$$3Cl_2 + 6OH^- \longrightarrow 5Cl^- + ClO_3^- + 3H_2O$$

Your notes

Your notes

Chemical Formulae & Oxidation Numbers

- Oxidation numbers are a useful tool for naming compounds as some elements can exist with more than one oxidation number
- For compound with two elements it is straight forward to name the compound
- For example
  - PCl_3 is phosphorus(III) chloride or phosphorus trichloride
  - PCl_5 is phosphorus(V) chloride or phosphorus pentachloride
  - OF_2 is oxygen difluoride
  - O_2F_2 is dioxygen difluoride
- In order to name a more complete compound we use Roman numerals for the element that has a variable oxidation number
  - K_2CrO_4 potassium chromate(VI)

Worked Example

Can you name these transition metal compounds?

1. Cu_2O
2. $MnSO_4$
3. Na_2CrO_4
4. $KMnO_4$
5. $Na_2Cr_2O_7$

Answer:

Answer 1: copper(I) oxide:

The ox. no. of 1 O atom is -2 and Cu_2O has overall no charge so the ox. no. of Cu is +1

Answer 2: manganese(II) sulfate:

The charge on the sulfate ion is -2, so the charge on Mn and ox. no. is +2

Answer 3: sodium chromate(VI):

The ox. no. of 2 Na atoms is +2 so CrO_4 has an overall -2 charge, so the ox. no. of Cr is +6

Answer 4: potassium manganate(VII):

The ox. no. of a K atom is +1 so MnO_4 has overall -1 charge, so the ox. no. of Mn is +7

Answer 5: sodium dichromate(VI):

The ox. no. of 2 Na atoms is +2 so Cr_2O_7 has an overall -2 charge, so the ox. no. of Cr is +6. To distinguish it from CrO_4 we use the prefix di in front of the anion

Your notes

Forming Anions & Cations

Metals

- Metals, in general, will form positive ions by losing electrons
- Therefore, they are **oxidised** and the oxidation number **increases**
- Example 1:
  - When sodium reacts with water, sodium hydroxide and hydrogen gas is formed

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

- The oxidation number of sodium changes from 0 to +1
- Example 2:
  - When magnesium reacts with hydrochloric acid, magnesium chloride and hydrogen gas is formed

$$Mg\ (s) + 2HCl\ (l) \rightarrow MgCl_2\ (aq) + H_2\ (g)$$

- The oxidation number of magnesium changed from 0 to +2

Non-metals

- Non-metals, in general, will form negative ions by gaining electrons
- Therefore, they are **reduced** and the oxidation number **decreases**
- Example:
  - When sodium reacts with oxygen, sodium oxide is formed

$$4Na\ (s) + O_2\ (g) \rightarrow Na_2O\ (s)$$

- The oxidation number of oxygen changes from 0 to -2

Constructing Ionic Equations

Half Equations and Ionic Equations

- Half equations and ionic equations are specific types of equations for showing some of the fine details going on in chemical reactions
- **Half equations** are used to show what happens to the electrons in reactions where atoms, molecules or ions are gaining or losing electrons
- They are called half equations, because they represent only *half* of what is happening in a reaction involving electron transfer
  - One species gains electrons

- Another species loses electrons

- Examples of half equations are:

$$Pb^{2+} + 2e^- \rightarrow Pb$$

$$2Br^- \rightarrow Br_2 + 2e^-$$

- Some half equations are more complicated and require the addition of water and hydrogen ions in addition to electrons
- The steps required to balance a half equation are:
    - **Step 1**: Write the unbalanced equation to show the species that undergoes reduction or oxidation, if necessary, balance the atom that is being reduced or oxidised
    - **Step 2**: Add H_2O to balance O atoms
    - **Step 3**: Add H^+ to balance H atoms
    - **Step 4**: Add e^- to balance the charge
- Half equations can be combined to form an **ionic equation**
- An **ionic equation** shows what happens in terms of ions in a chemical reaction
- This is best shown with an example:

Worked Example

Dichromate ions, $Cr_2O_7^{2-}$ ions react with iron(II) ions, Fe^{2+}, in acidic conditions, forming chromium(III) ions, Cr^{3+}, and iron(III) ions, Fe^{3+}. Write the half equations and use these to construct an ionic equation for this reaction.

Answer:

- Fe^{2+} ions are being oxidised as they gain an electron to form Fe^{3+}, this half equation can be simply constructed showing the loss of an electron:
    - $Fe^{2+} \rightarrow Fe^{3+} + e^-$
- The reduction of $Cr_2O_7^{2-}$ ions to Cr^{3+} is not as straightforward, so use the steps above to construct the half equation:
    - **Step 1**: Write the unbalanced equation to show the species that undergoes reduction or oxidation, if necessary, balance the atom that is being reduced or oxidised:
        - $Cr_2O_7^{2-} \rightarrow \mathbf{2}Cr^{3+}$
    - **Step 2**: Add H_2O to balance O atoms
        - $Cr_2O_7^{2-} \rightarrow 2Cr^{3+} + \mathbf{7H_2O}$
    - **Step 3**: Add H^+ to balance H atoms
        - $Cr_2O_7^{2-} + \mathbf{14H^+} \rightarrow 2Cr^{3+} + 7H_2O$

Your notes

- **Step 4**: Add e⁻ to balance the charge
  - $Cr_2O_7^{2-} + 14H^+ + \mathbf{6e^-} \rightarrow 2Cr^{3+} + 7H_2O$
- To construct the full ionic equation, combine the two half equations
  - To do this, the number of electrons in each half equation needs to be the same so that they cancel out when the equations are combined, and so that electrons do not appear in the ionic equation
  - In this example, multiply the first half equation by 6:
    - $\mathbf{6}Fe^{2+} \rightarrow \mathbf{6}Fe^{3+} + \mathbf{6}e^-$
    - $Cr_2O_7^{2-} + 14H^+ + 6e^- \rightarrow 2Cr^{3+} + 7H_2O$
  - Now add all of the reactants and products together - the electrons will cancel to give the full ionic equation:
    - $6Fe^{2+} + Cr_2O_7^{2-} + 14H^+ \rightarrow 6Fe^{3+} + 2Cr^{3+} + 7H_2O$

Your notes

Balancing full ionic equations

- You do not have to construct the half equations to be able to write a full ionic equation, you can use oxidation numbers to balance them
- This is not as straightforward as just balancing the atoms involved
- Balancing equations using redox principles is a useful skill and is best illustrated by following an example
- It is important to follow a methodical step-by-step approach so that you don't get lost:

Worked Example

Writing overall redox reactions

Manganate(VII) ions (MnO_4^-) react with Fe^{2+} ions in the presence of acid (H^+) to form Mn^{2+} ions, Fe^{3+} ions and water

Write the overall redox equation for this reaction

Answer

Step 1: Write the unbalanced equation and identify the atoms which change in oxidation number

$$\underset{+7\ -2}{MnO_4^-}(aq) + \underset{+2}{Fe^{2+}}(aq) + \underset{+1}{H^+}(aq) \longrightarrow \underset{+2}{Mn^{2+}}(aq) + \underset{+3}{Fe^{3+}}(aq) + \underset{+1\ -2}{H_2O}(l)$$

Step 2: Deduce the oxidation number changes

Your notes

OX. NO. CHANGE = −5

$$\underset{+7}{MnO_4^-}(aq) + \underset{+2}{Fe^{2+}}(aq) + H^+(aq) \longrightarrow \underset{+2}{Mn^{2+}}(aq) + \underset{+3}{Fe^{3+}}(aq) + H_2O(l)$$

OX. NO. CHANGE = +1

Step 3: Balance the oxidation number changes

OX. NO. CHANGE = 1 × (−5) = −5

$$\underset{+7}{MnO_4^-}(aq) + 5\underset{+2}{Fe^{2+}}(aq) + H^+(aq) \longrightarrow \underset{+2}{Mn^{2+}}(aq) + 5\underset{+3}{Fe^{3+}}(aq) + H_2O(l)$$

OX. NO. CHANGE = 5 × (+1) = +5

Step 4: Balance the charges

$$MnO_4^-(aq) + 5Fe^{2+}(aq) + H^+(aq) \longrightarrow Mn^{2+}(aq) + 5Fe^{3+}(aq) + H_2O(l)$$

IGNORING H^+

- TOTAL ⊕ CHARGE = (5 × 2+) = 10+
- TOTAL ⊖ CHARGE = 1−

TOTAL CHARGE REACTANTS = 9+

- TOTAL ⊕ CHARGE = 5 × (3+) + (2+) = 17+
- TOTAL ⊖ CHARGE = 0

TOTAL CHARGE PRODUCTS = 17+

THUS $8H^+$ IONS ARE NEEDED TO BALANCE THE CHARGES ON BOTH SIDES

Step 5: Finally, balance the atoms

$$MnO_4^-(aq) + 5Fe^{2+}(aq) + 8H^+(aq) \longrightarrow Mn^{2+}(aq) + 5Fe^{3+}(aq) + 4H_2O(l)$$

Acid-Base Titrations with Indicators

- **Acid-base titrations** are used to find the unknown concentrations of solutions of acids and bases
- **Acid-base indicators** give information about the change in chemical environment
- They change colour reversibly depending on the concentration of H^+ ions in the solution
- **Indicators** are weak acids and bases where the conjugate bases and acids have a different colour
- Many **acid-base indicators** are derived from plants, such as litmus

Common Indicators Table

Indicator	Colour in acid	Colour in alkali
Litmus	red	blue
Methyl orange	red	yellow
Phenolphthalein	colourless	pink

- A good indicator gives a very sharp colour change at the **equivalence point**
- In **titrations** is it not always possible to use two colour indicators because of this limitation, so for example litmus cannot be used successfully in a **titration**
- When **phenolphthalein** is used, it is usually better to have the base in the burette because it is easier to see the sudden and permanent appearance of a colour (pink in this case) than the change from a coloured solution to a colourless one

Volumes & concentrations of solutions

- The **concentration** of a solution is the amount of **solute** dissolved in a **solvent** to make 1 dm^3 of **solution**
  - The solute is the substance that dissolves in a solvent to form a solution
  - The solvent is often water

$$\text{Concentration (mol dm}^{3}\text{)} = \frac{\text{number of moles of solute (mol)}}{\text{volume of solution (dm}^3\text{)}}$$

- A **concentrated** solution is a solution that has a **high** concentration of solute
- A **dilute** solution is a solution with a **low** concentration of solute

- When carrying out calculations involve concentrations in mol dm^{-3} the following points need to be considered:
  - Change mass in grams to **moles**
  - Change cm^3 to dm^3
- To calculate the **mass** of a substance present in solution of known **concentration and volume**:
  - Rearrange the concentration equation

number of moles (mol) = concentration (mol dm^{-3}) x volume (dm^3)

- Multiply the moles of solute by its molar mass

mass of solute (g) = number of moles (mol) x molar mass (g mol^{-1})

Worked Example

Neutralisation calculation

25.0 cm^3 of 0.050 dm^{-3} sodium carbonate was completely neutralised by 20.00 cm^3 of dilute hydrochloric acid. Calculate the concentration in mol dm^{-3} of the hydrochloric acid.

Answer

Step 1: Write the balanced symbol equation

$$Na_2CO_3 + 2HCl \rightarrow 2NaCl + H_2O + CO_2$$

Step 2: Calculate the amount, in moles, of sodium carbonate reacted by rearranging the equation for amount of substance (mol) and dividing the volume by 1000 to convert cm^3 to dm^3

amount (Na_2CO_3) = 0.025 dm^3 x 0.050 mol dm^{-3} = 0.00125 mol

Step 3: Calculate the moles of hydrochloric acid required using the reaction's stoichiometry

1 mol of Na_2CO_3 reacts with 2 mol of HCl, so the molar ratio is 1 : 2

Therefore 0.00125 moles of Na_2CO_3 react with 0.00250 moles of HCl

Step 4: Calculate the concentration, in mol dm^{-3}, of hydrochloric acid

$$\text{concentration (HCl) (mol dm}^{-3}\text{)} = \frac{\text{amount (mol)}}{\text{volume (dm}^3\text{)}}$$

$$\text{concentration (HCl) (mol dm}^{-3}\text{)} = \frac{0.00250}{0.0200}$$

Your notes

concentration (HCl) (mol dm^{-3}) = 0.125 mol dm^{-3}

Your notes

Worked Example

Concentration in g dm^{-3}

A student dissolved 10 g of sodium hydroxide, NaOH, in 2 dm^3 of distilled water. Calculate the concentration of the solution.

Answer:

- $\text{Concentration} = \dfrac{\text{Mass (g)}}{\text{Volume (dm}^3\text{)}}$
- $\text{Concentration} = \dfrac{10}{2} = 5 \text{ g dm}^{-3}$

Uncertainty – Calculations

Percentage Uncertainties

- **Percentage uncertainties** are a way to compare the significance of an **absolute uncertainty** on a measurement
- This is not to be confused with **percentage error**, which is a comparison of a result to a literature value
- The formula for calculating percentage uncertainty is as follows:

$$\text{percentage uncertainty (\%)} = \frac{\text{total uncertainty}}{\text{measured value}} \times 100$$

Adding or subtracting measurements

- When you are adding or subtracting two measurements then you add together the **absolute** measurement uncertainties
- For example,
  - Using a balance to measure the initial and final mass of a container
  - Using a thermometer for the measurement of the temperature at the start and the end
  - Using a burette to find the initial reading and final reading
- In all of these examples, you have to read the instrument **twice** to obtain the quantity
- If each time you read the instrument the measurement is 'out' by the stated uncertainty, then your final quantity is potentially 'out' by **twice** the uncertainty

Total experimental uncertainty

Your notes

- Some experiments use multiple pieces of equipment, which each have a measure of uncertainty in their use
- For example, a titration against a standard solution will have the following uncertainties:
  - Using a balance to measure the mass of the solid used to make the standard solution
  - Using a volumetric flask to make the standard solution
  - Using a volumetric pipette to measure the sample that is being titrated against
  - Using a burette to find the initial reading and final reading that determine the titre volume
- If each piece of equipment used in the experiment has its own associated uncertainty, you can determine the total experimental uncertainty by adding the total individual uncertainties

$$\text{total experimental percentage uncertainty } (\%) = \Sigma(\text{individual uncertainties })$$

- For example, a titration against a standard solution may have:
  - An uncertainty of ±0.5% for the use of the balance
  - An uncertainty of ±0.1% for the use of the volumetric flask
  - An uncertainty of ±0.2% for the use of the volumetric pipette
  - An uncertainty of ±0.4% for the use of the burette
  - Therefore, the total experimental uncertainty would be 0.5 + 0.1 + 0.2 + 0.4 = ±1.2%