MODULE 2

Foundations in Chemistry

Chapters in this Module

Demonstrate knowledge, understanding, and application of:

1. Atoms, ions, and compounds
2. Amount of substance
3. Acids and redox
4. Electrons and bonding
5. Shapes of molecules and intermolecular forces

Introduction

Chemistry is about all matter and the chemical reactions that take place between atoms. This makes chemistry a vast subject, whether you are studying the 34 elements making up a human body, synthesising a new medicine or inventing nanoparticles for a use yet to be found. This module studies the foundations of chemistry that form the building blocks for all the other modules in your A level course.

Atoms, ions, and compounds looks at some of the essential language of chemistry. You will learn about the atomic masses that you see on your periodic table. You will also learn about the special code of chemistry: the formulae, and equations that allow chemists to communicate.

Amount of substance and its unit the mole, provides chemists with an ability to convert between mass, concentration and volume to predict how much product can be made in a chemical reaction.

Acids and redox are two important topics. Central to acids is the analysis of solutions by titration, using pipettes and burettes, the basis of quality testing for washing powders, and medicines. In redox, you learn about oxidation numbers another essential part of the chemists toolkit for describing chemical change.

Electrons and bonding looks at the role of electrons in atoms and in chemical bonding. A good understanding of bonding and structure is essential for all further topics.

Shapes of molecules and intermolecular forces is all about the molecule. You will see how electrons determine the shape and polarity of molecule. You will learn about how intermolecular forces explain many properties of molecular compounds such as why ice floats and why water is a liquid.

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Knowledge and understanding checklist

From your Key Stage 4 study you should be able to answer the following questions. Work through each point, using your Key Stage 4 notes and the support available on Kerboodle.

- [ ] recall relative charges and approximate relative masses of protons, neutrons and electrons
- [ ] calculate numbers of protons, neutrons and electrons in atoms, given atomic number and mass number
- [ ] write formulae and balanced chemical equations
- [ ] calculate relative formula masses of species separately and in a balanced chemical equation
- [ ] use a balanced equation to calculate masses of reactants or products
- [ ] recall that acids react with some metals and with carbonates and write equations predicting products from given reactants
- [ ] describe neutralisation as acid reacting with alkali to form a salt plus water
- [ ] explain reduction and oxidation in terms of gain or loss of electrons, identifying which species are oxidised and which are reduced.
- [ ] construct dot and cross diagrams for simple ionic and covalent substances

Maths skills checklist

In this unit, you will need to use the following maths skills. You can find support for these skills on Kerboodle and through MyMaths.

- [ ] **Working with standard form and significant figures, and using appropriate units**. You will need this when carrying out all calculations in this chapter.
- [ ] **Changing the subject of an equation**. You will need this when carrying out structured and unstructured mole calculations.
- [ ] **Using ratios, fractions and percentages**. You will need this when working with moles and equations using ratios, calculating percentage yields, and calculating atom economies.
- [ ] **Finding arithmetic means**. You will need this when calculating weighted means when determining an atomic mass and when calculating mean titres.
- [ ] **Using angles and shapes in regular 2-D and 3-D structures**. You will need this when predicting the shapes of and bond angles in molecule and ions.

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4 ACIDS AND REDOX

4.1 Acids, bases, and neutralisation

Specification reference: 2.1.4

Learning outcomes

Demonstrate knowledge, understanding, and application of:

- acids and bases
- neutralisation.

▼ **Table 1** *Common acids*

Acid	Formula
hydrochloric acid	HCl
sulfuric acid	H_2SO_4
nitric acid	HNO_3
ethanoic acid (vinegar)	CH_3COOH

Synoptic link

You will learn more about equilibrium and weak acids in Chapter 20.



▲ **Figure 1** *Some common acids – vinegar contains ethanoic acid, orange and lemon juice and even sink cleaner contains citric acid, and apples contain malic acid*

Acids

All acids contain hydrogen in their formulae (Table 1). When dissolved in water, an acid releases hydrogen ions as protons, H^+, into the solution. In the equation below, hydrogen chloride gas releases H^+ ions as it dissolves in water.

$$HCl(g) + aq \rightarrow H^+(aq) + Cl^-(aq)$$

In this equation + aq has been included to show that an excess of water is present. The equation is essentially hydrogen chloride gas dissolving to form an aqueous solution.

Strong and weak acids

A **strong acid**, such as hydrochloric acid, HCl, releases all its hydrogen atoms into solution as H^+ ions and **completely dissociates** in aqueous solution.

$$HCl(aq) \rightarrow H^+(aq) + Cl^-(aq)$$

A **weak acid**, such as ethanoic acid, CH_3COOH, only releases a small proportion of its available hydrogen atoms into solution as H^+ ions. A weak acid **partially dissociates** in aqueous solution.

$$CH_3COOH(aq) \rightleftharpoons H^+(aq) + CH_3COO^-(aq)$$

The equilibrium sign $\rightleftharpoons$ indicates that the forward reaction is incomplete.

It is important to realise that not all compounds that contain hydrogen atoms are acids. Each molecule of ethanoic acid contains four hydrogen atoms, but only the hydrogen atom on the COOH group is released as H^+. Even then, only about one molecule in every hundred dissociates, so ethanoic acid is a weak acid. Most organic acids, like ethanoic acid, are weak acids.

Bases and alkalis

Metal oxides, metal hydroxides, metal carbonates, and ammonia, NH_3, are classified as **bases**. A base **neutralises** an acid to form a **salt**. Table 2 and Figure 2 show some common bases.

An **alkali** is a base that dissolves in water releasing hydroxide ions (OH^-) into the solution. The equation below shows the alkali sodium hydroxide releasing hydroxide ions as it dissolves in water.

$$NaOH(s) + aq \rightarrow Na^+(aq) + OH^-(aq)$$

▼ **Table 2** *Common bases*

Metal oxides	Metal carbonates	Alkalis
MgO	Na_2CO_3	NaOH
CaO	$CaCO_3$	KOH
CuO	$CuCO_3$	NH_3

Neutralisation

In neutralisation of an acid, $H^+(aq)$ ions react with a base to form a salt and neutral water. The H^+ ions from the acid are replaced by metal or ammonium ions from the base. Table 3 shows salts of common acids.

- Notice the link between the name of the acid and the salt, in blue.
- To form the salt, the hydrogen (shown in purple) in the acid is replaced by a metal or ammonium ion to form the salt.

▼ **Table 3** *Acids and their salts*

Acid		Salt		
Name	Formula	Type	Name	Formula
hydrochloric acid	HCl	chloride	sodium chloride	NaCl
sulfuric acid	H_2SO_4	sulfate	sodium sulfate	Na_2SO_4
nitric acid	HNO_3	nitrate	calcium nitrate	$Ca(NO_3)_2$
ethanoic acid (vinegar)	CH_3COOH	ethanoate	potassium ethanoate	CH_3COOK

Neutralisation of acids with metal oxides and hydroxides

An acid is neutralised by a metal oxide or metal hydroxide to form a salt and water only. The equations below show the neutralisation of sulfuric acid and hydrochloric acid by copper(II) oxide to form a salt and water only. Figure 3 shows solutions of the salts formed.

$$CuO(s) + H_2SO_4(aq) \rightarrow CuSO_4(aq) + H_2O(l)$$
$$CuO(s) + 2HCl(aq) \rightarrow CuCl_2(aq) + H_2O(l)$$

Alkalis

With alkalis, the reactants are in solution. As with metal oxides, the overall reaction forms a salt and water only:

$$acid + alkali \rightarrow salt + water$$

The ionic equation, shown below, is much simpler than the overall equation: neutralisation of $H^+(aq)$ ions by $OH^-(aq)$ ions to form neutral water, $H_2O(l)$.

Full equation: $HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)$

Ionic equation: $H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$



▲ **Figure 2** *Three bases – calcium carbonate, $CaCO_3$, copper oxide, CuO, and sodium hydroxide, NaOH*

Study tip

The reactions for the neutralisation of an acid by a metal oxide or hydroxide are all essentially the same:

acid + metal oxide/hydroxide → salt + water

Learn this to help you write the equation for any neutralisation of this type.



▲ **Figure 3** *Solutions of two copper salts prepared by neutralisation of two different acids with the base copper(II) oxide:*

- *sulfuric acid has been neutralised to form blue copper(II) sulfate, $CuSO_4$ (left)*
- *hydrochloric acid has been neutralised to form green copper(II) chloride, $CuCl_2$ (right)*

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Study tip

Remember the equation for the neutralisation of acids by metal carbonates.

acid + metal carbonate → salt + water + carbon dioxide(g)

It will help you write the equation for any neutralisation of this type.

Neutralisation of acids with carbonates

Like metal oxides, carbonates neutralise acids to form a salt and water, but also a third product, carbon dioxide gas. The equations show neutralisation of two carbonates by sulfuric acid and hydrochloric acid.

$$ZnCO_3(s) + H_2SO_4(aq) \rightarrow ZnSO_4(aq) + H_2O(l) + CO_2(g)$$

$$MgCO_3(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2O(l) + CO_2(g)$$

Dissociation in sulfuric acid

Sulfuric acid, H_2SO_4, is a strong acid, but this is true only for one of the two hydrogen atoms. When sulfuric acid is mixed with water each H_2SO_4 molecule dissociates, releasing just one of its two hydrogen atoms as an H^+ ion:

$$H_2SO_4(aq) \rightarrow H^+(aq) + HSO_4^-(aq)$$

The resulting $HSO_4^-(aq)$ ions then only partially dissociate:

$$HSO_4^-(aq) \rightleftharpoons H^+(aq) + SO_4^{2-}(aq)$$

- Sulfuric acid first behaves as a strong acid.
- The HSO_4^- ions formed behave as a weak acid.

Other strong acids containing more than one hydrogen atom behave similarly.

Write equations to show the dissociation of the three hydrogen atoms in phosphoric acid, H_3PO_4.

Summary questions

1 Explain what is meant by the following terms.
- **a** strong acid *(1 mark)*
- **b** weak acid *(1 mark)*

2 Write equations to show the dissociation of the following acids when dissolved in water.
- **a** nitric acid, HNO_3 (a strong acid) *(2 marks)*
- **b** propanoic acid, CH_3CH_2COOH (a weak acid) *(2 marks)*

3 Write equations for the following neutralisation reactions. For each reaction, name the salt formed.
- **a** The reaction of $MgO(s)$ with hydrochloric acid *(2 marks)*
- **b** The reaction of $NaOH(aq)$ with sulfuric acid *(2 marks)*
- **c** The reaction of $ZnCO_3(s)$ with nitric acid *(2 marks)*
- **d** The reaction of aqueous sodium hydroxide with ethanoic acid *(2 marks)*

4 Baking powder usually contains an organic acid and sodium hydrogencarbonate, $NaHCO_3$, also known as bicarbonate of soda. Sodium hydrogencarbonate is an acid salt, formed by partial neutralisation of carbonic acid, H_2CO_3, with sodium hydroxide.
- **a** Write equations for the partial and complete neutralisation reactions of carbonic acid with sodium hydroxide. *(4 marks)*
- **b** Why is sodium hydrogencarbonate called an acid salt? *(1 mark)*
- **c** Baking powder is used to make cakes. When it is mixed into a batter or dough containing water and heated, the sodium hydrogencarbonate reacts with the acid in the powder in the same way as a carbonate. The small quantity of baking powder does not affect the taste, so why is it used? *(2 marks)*

4.2 Acid–base titrations

Specification reference: 2.1.4



Learning outcomes

Demonstrate knowledge, understanding, and application of:
- preparation of a standard solution
- carrying out a titration
- analysing titration results by calculation.

Titrations

A **titration** is a technique used to accurately measure the volume of one solution that reacts exactly with another solution. Titrations can be used for:

- finding the concentration of a solution
- identification of unknown chemicals
- finding the purity of a substance.

Checking purity is an important aspect of quality control, especially for compounds manufactured for human use such as medicines, food, and cosmetics. It is essential that pharmaceuticals have a high level of purity – just a tiny amount of an impurity in a drug could cause a great deal of harm to a patient.

Synoptic link

Standard solutions were introduced in Topic 3.3, Moles and volumes.

Preparing a standard solution

A standard solution is a solution of known concentration.
A **volumetric flask** is used to make up a standard solution very accurately. Volumetric flasks are manufactured in various sizes and can measure volumes very precisely. The volumetric flasks that you will use are manufactured to the typical tolerances below:

- a 100 cm^3 volumetric flask: ±0.20 cm^3
- a 250 cm^3 volumetric flask: ±0.30 cm^3.

Preparing standard solutions

1. The solid is first weighed accurately.
2. The solid is dissolved in a beaker using less distilled water than will be needed to fill the volumetric flask to the mark.
3. This solution is transferred to a volumetric flask. The last traces of the solution are rinsed into the flask with distilled water.
4. The flask is carefully filled to the graduation line by adding distilled water a drop at a time until the *bottom* of the meniscus lines up exactly with the mark (Figure 1). Care at this stage is essential – if too much water is added, the solution will be too dilute and must be prepared again. You should view the graduation mark and meniscus at eye level for accuracy.
5. Finally, the volumetric flask is slowly inverted several times to mix the solution thoroughly. If this stage is omitted, titration results are unlikely to be consistent. You will be able to see the solution mixing when you invert the flask as the more dense original solution moves through the solution.

Explain the effect on the titre of the following errors.
1. The flask is filled with water above the graduation line.
2. The flask is not inverted.



▲ **Figure 1** *The volumetric flask is filled so that bottom of the meniscus just touches the graduation line*



▲ **Figure 2** *The volumetric flask is slowly inverted several times to ensure that the solution is mixed evenly*

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Acid–base titrations

Apparatus

In an acid–base titration, a solution of an acid is titrated against a solution of a base using a **pipette** and a **burette**, which are typically manufactured to the tolerances below:

- a 10 cm^3 pipette: ±0.04 cm^3
- a 25 cm^3 pipette: ±0.06 cm^3
- a 50 cm^3 burette: ±0.10 cm^3.

A burette reading is recorded to the nearest half division, with the bottom of the meniscus on a mark or between two marks. Each burette reading is measured to the nearest ±0.05 cm^3 so the reading always has *two* decimal places, the last place being either 0 or 5, for example, 25.40 cm^3 or 26.25 cm^3.



▲ **Figure 3** *When filling a burette, run excess solution out through the tap to remove any air bubbles. If a bubble is left in the neck of the burette, the air could be released during the titration, leading to an error in the titre*



▲ **Figure 4** *Take burette readings from the bottom of the meniscus, with your eye is at the level of meniscus. Read the burette to the nearest 0.05 cm^3. This burette reading is 25.55 cm^3*

The acid–base titration procedure

1. Add a measured volume of one solution to a conical flask using a pipette.
2. Add the other solution to a burette, and record the initial burette reading to the nearest 0.05 cm^3.
3. Add a few drops of an **indicator** to the solution in the conical flask.
4. Run the solution in the burette into the solution in the conical flask, swirling the conical flask throughout to mix the two solutions. Eventually the indicator changes colour at the **end point** of the titration. The end point is used to indicate the volume of one solution that exactly reacts with the volume of the second solution.
5. Record the final burette reading. The volume of solution added from the burette is called the titre, which is calculated by subtracting the initial from the final burette reading.
6. A quick, trial titration is carried out first to find the approximate **titre**.
7. The titration is then repeated accurately, adding the solution dropwise as the end point is approached. Further titrations are carried out until two accurate titres are **concordant** – agreeing to within 0.10 cm^3.

The readings from the titration are recorded in a table such as Table 1.

▼ **Table 1** *Table for recording titration results*

	Trial	1	2	3
final burette reading / cm^3				
initial burette reading / cm^3				
titre / cm^3				
mean titre / cm^3				

Explain the effect on the titre of the following errors.

1. The pipette has an air bubble inside.
2. The burette readings are taken from the top, rather than the bottom, of the meniscus.

The mean titre

When working out the mean titre, it is important to use only your *closest accurate titres*.

- By repeating titres until two agree within 0.10 cm^3, you can reject inaccurate titres.
- If you were to include all the titres in the mean, you have lost the accuracy of the titration technique.

Titration calculations

From the results of a titration, you will know the following:

- *both* the concentration c_1 and the reacting volume V_1 of one of the solutions
- *only* the reacting volume V_2 of the other solution.

The method for analysing the results follows a set pattern.

Step 1: Work out the amount, in mol, of the solute in the solution for which you know *both* the concentration c_1 and volume V_1.

Step 2: Use the equation to work out the amount, in mol, of the solute in the other solution.

Step 3: Work out the unknown information about the solute in the other solution.

Worked example: Determination of an unknown concentration

Pipette: 25.00 cm^3 of 0.100 mol dm^{-3} KOH(aq)

Mean titre from burette: 25.70 cm^3 of H_2SO_4(aq)

Unknown information: The concentration of H_2SO_4(aq)

Step 1: From the titration results, calculate the amount of KOH.

$$n(\text{KOH}) = c \times \frac{V}{1000} = 0.100 \times \frac{25.00}{1000} = \mathbf{0.002\,50}\ \text{mol}$$

Step 2: From the equation and Step 1, determine the amount of H_2SO_4.

$$\mathbf{2}\ \text{KOH(aq)} + \text{H}_2\text{SO}_4\text{(aq)} \rightarrow \text{K}_2\text{SO}_4\text{(aq)} + 2\ \text{H}_2\text{O(l)}$$

2 mol **1** mol (balancing numbers)

0.002 50 mol 0.00125 mol

Step 3: Work out the unknown information.

$$n(\text{H}_2\text{SO}_4) = c \times \frac{V(\text{cm}^3)}{1000}$$

$$c = \frac{1000 \times n}{V} = \frac{1000 \times 0.00125}{25.70} = 0.0486\ \text{mol dm}^{-3}$$

concentration of H_2SO_4(aq) is 0.0486 mol dm^{-3}

Study tip

This method is essentially the same as the one used to calculate unknown quantities using the mole in Topic 3.4, Reacting quantities.

Study tip

It's all done with the ratios of the balancing numbers and moles.

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Identification of a carbonate

You can use the titration technique to work out some unknown information about a substance, for example, you can identify an unknown carbonate, $\mathbf{X}_2CO_3$.

The key steps are shown below, together with results.

1 Prepare a solution of an unknown carbonate, $\mathbf{X}_2CO_3$, in a volumetric flask.
2 Using a pipette, measure 25.00 cm^3 of your prepared solution into a conical flask.
3 Using a burette, titrate this solution using 0.100 mol dm^{-3} hydrochloric acid.
4 Analyse your results to identify the carbonate.

Mass measurements

mass of weighing bottle / g	11.41
mass of weighing bottle + $\mathbf{X}_2CO_3$ / g	12.60
mass of $\mathbf{X}_2CO_3$ / g	1.19

Titration readings

	Trial	1	2	3
final burette reading / cm^3	24.10	22.35	44.95	22.45
initial burette reading / cm^3	1.00	0.00	22.35	0.00
titre / cm^3	23.10	**22.35**	22.60	**22.45**
mean titre / cm^3			**22.40**	

Analysis

Step 1: Calculate the amount of HCl that reacted.

Use the mean titre V and the concentration of the hydrochloric acid c.

$$n(HCl) = c \times \frac{V}{1000} = 0.100 \times \frac{22.40}{1000} = \mathbf{0.002\,24\ mol}$$

Step 2: Determine the amount of $\mathbf{X}_2CO_3$ that reacted.

Use the equation and n(HCl).

$\mathbf{X}_2CO_3(aq)$ + **2** HCl(aq) → $X_2SO_4(aq) + H_2O(l) + CO_2(g)$

1 mol **2** mol (balancing numbers)

0.001 12 mol **0.002 24 mol HCl**

Step 3: Work out the unknown information. There are several stages.

1 Scale up to find the amount of $\mathbf{X}_2CO_3$ in the 250 cm^3 solution that you prepared.
$n(X_2CO_3)$ in 25.00 cm^3 used in the titration = 0.001 12 mol
$n(X_2CO_3)$ in 250.0 cm^3 solution = 0.001 12 × 10 = **0.0112** mol

2 Find the molar mass of $\mathbf{X}_2CO_3$.
Use the amount, $n(X_2CO_3)$, in the 250 cm^3 solution and the mass m of $\mathbf{X}_2CO_3$ used to prepare this solution.

$$n(\mathbf{X}_2CO_3) = \frac{m}{M}$$

$$M(\mathbf{X}_2CO_3) = \frac{m}{N} = \frac{1.19}{0.0112} = \mathbf{106.25}\ g\,mol^{-1}$$

3 Finally use $M(\mathbf{X}_2CO_3)$ to identify **X** in the formula $\mathbf{X}_2CO_3$.

$$106.25 = M(\mathbf{X}) \times 2 + 12.0 + (16.0 \times 3) = 2A_r + 60$$

$$M(\mathbf{X}) = \frac{106.25 - 60}{2} = \mathbf{23.125}\ g\,mol^{-1}$$

From the periodic table, 23.125 most closely matches Na (A_r = 23.0).

Unknown carbonate is sodium carbonate, Na_2CO_3.

Use the raw results below to determine the molar mass of an unknown acid HA.

Mass readings: For preparation of a 250.0 cm^3 solution

mass of weighing bottle / g	12.51
mass of weighing bottle + HA / g	9.64

Titration readings: Titration with 0.0600 mol dm^{-3} Na_2CO_3

	Trial	1	2	3
final burette reading / cm^3	27.40	25.05	29.55	24.65
initial burette reading / cm^3	2.00	0.00	5.00	0.00

Study tip

The scaling here is obvious because 250 is 10 times larger than 25. If you were scaling up a burette reading (say 29.55 cm^3), the process is the same, but you would need to find the scaling factor by dividing 250 by the titre.

Study tip

You might be given a *structured calculation* – you will be helped through the calculation with prompts similar to the bullet points in this example.

But you might also be given an *unstructured calculation* with no help. You should practise both structured titration calculations and the harder unstructured problems.

Summary questions

1 a 25.00 cm^3 of 0.110 mol dm^{-3} NaOH(aq) reacts exactly with 23.30 cm^3 of HNO_3(aq). *(5 marks)*
i Calculate the amount, in moles, of NaOH(aq) used.
ii Write the equation for the reaction and calculate the amount of HNO_3(aq) in the titre.
iii Calculate the concentration of the HNO_3(aq) solution.
b 25.00 cm^3 of 0.125 mol dm^{-3} KOH(aq) reacts exactly with 26.60 cm^3 of H_2SO_4(aq). Write the equation and find the unknown concentration. *(3 marks)*

2 1.96 g of an unknown hydroxide $X(OH)_2$ was dissolved in water, and the solution was made up to 250.00 cm^3 in a volumetric flask. In a titration, 25.00 cm^3 of this solution of $X(OH)_2$ required 21.20 cm^3 of 0.250 mol dm^{-3} HCl(aq) to reach the end point.

Equation: $2HCl(aq) + X(OH)_2(aq) \rightarrow XCl_2(aq) + 2H_2O(l)$
Identify the unknown hydroxide $X(OH)_2$. *(4 marks)*

3 1.654 g of a hydrated acid $H_2C_2O_4 \cdot xH_2O$ was dissolved in water and the solution was made up to 250.00 cm^3 in a volumetric flask. 23.80 cm^3 of this solution required 25.00 cm^3 of 0.100 mol dm^{-3} NaOH(aq) to reach the end point.

Equation: $H_2C_2O_4(aq) + 2NaOH(aq) \rightarrow Na_2C_2O_4(aq) + 2H_2O(l)$
Identify the value of x and hence the formula of the hydrated acid $H_2C_2O_4 \cdot xH_2O$. *(4 marks)*

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4.3 Redox

Specification reference: 2.1.5

Learning outcomes

Demonstrate knowledge, understanding, and application of:

- oxidation number
- oxidation and reduction
- redox reactions.

Study tip

It is a common mistake to confuse oxidation numbers of elements and ions. The oxidation number of oxygen in O_2 is zero but in H_2O is –2.

Oxidation number

Oxidation number is based on a set of rules that apply to atoms, and can be thought of as the number of electrons involved in bonding to a different element. Use of oxidation numbers helps when writing formulae and balancing electrons as a check that all electrons have been accounted for.

Rules for elements

The oxidation number is *always* zero for elements.

- In a pure element, any bonding is to atoms of the same element.
- So in H_2, O_2, P_4, S_8, Na, and Fe the oxidation number of each atom of the element is 0.

Rules for compounds and ions

- Each atom in a compound has an oxidation number.
- An oxidation number has a sign, which is placed *before* the number.

Table 1 shows examples of oxidation numbers of atoms in compounds and ions, including some special cases. The oxidation number of an ion of an element is numerically the same as the ionic charge but the sign comes before the number.

Study tip

All oxidation numbers, except zero (0), have a sign, + or –.

The *sign* of an oxidation number is placed *before* the number: the oxidation number of Ca in a Ca^{2+} ion is +2, *not* 2+.

▼ **Table 1** *Oxidation number rules in compounds and ions*

Combined element	Oxidation number	Examples
O	–2	H_2O, CaO.
H	+1	NH_3, H_2S.
F	–1	HF
Na^+, K^+	+1	NaCl, K_2O
Mg^{2+}, Ca^{2+}	+2	$MgCl_2$, CaO
Cl^-, Br^-, I^-	–1	HCl, KBr, CaI_2
Special cases:		
H in metal hydrides	–1	NaH, CaH_2
O in peroxides	–1	H_2O_2
O bonded to F	+2	F_2O

Working out oxidation numbers

In addition to the rules in Table 1, there is a further rule for combined atoms:

- sum of the oxidation numbers = total charge.

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Worked example: Oxidation number in compounds

What is the oxidation number of sulfur in sulfuric acid, H_2SO_4?

Step 1: Assign any oxidation numbers from the rules.

H_2SO_4

Total H = +1 × 2 = +2 { +1, +1 } Total O = –2 × 4 = –8 { –2, –2, –2, –2 }

Step 2: What is the sum of oxidation numbers?

sum of oxidation numbers = total charge = 0

Step 3: Work out the unknown oxidation numbers.

sum of oxidation numbers = (+2) + (X) + (–8) = 0

H_2SO_4

Oxidation number of sulfur in H_2SO_4 = +6

Study tip

Remember: the sum of the oxidation numbers must equal the overall charge.

Worked example: Oxidation numbers in ions

What is the oxidation number of nitrogen in NO_3^-?

Step 1: Assign any oxidation numbers from the rules.

NO_3^-

{ –2, –2, –2 } Total O = –2 × 3 = –6

Step 2: What is the sum of oxidation numbers?

sum of oxidation numbers = total charge = –1

Step 3: Work out the unknown oxidation numbers

sum of oxidation numbers = (X) + (–6) = –1

NO_3^-

Oxidation number of nitrogen in NO_3 = +5

Using Roman numerals in naming

Roman numerals are used in the names of compounds of elements that form ions with different charges. The Roman numeral shows the oxidation state (oxidation number) of the element, without a sign. The sign of the oxidation state is obvious from the overall charge:

- iron(II) represents Fe^{2+} with oxidation number +2
- iron(III) represents Fe^{3+} with oxidation number +3.

You have already seen that polyatomic ions containing oxygen, such as NO_2^- and NO_3^-, are sometimes named using –ite and –ate. Although still in common usage, use of –ite and –ate in naming is old-fashioned and modern names use oxidation numbers shown as a Roman numerals.

▼ **Table 2** *Naming of polyatomic ions*

Ion	Common name	Oxidation number of nitrogen	Modern name
NO_2^-	nitrite	+3	nitrate(III)
NO_3^-	nitrite	+5	nitrate(V)

In common usage, the Roman numeral is often omitted for the common ion, usually with more oxygen atoms:

- nitrate is assumed to be NO_3^-
- sulfate is assumed to be SO_4^{2-}.

Redox reactions

Reduction and oxidation

Originally the terms *oxidation* and *reduction* were used solely for reactions involving oxygen.

- Oxidation is addition of oxygen.
- Reduction is removal of oxygen.

The reaction below shows oxidation and reduction:

$$CuO(s) + H_2(g) \rightarrow Cu(s) + H_2O(l)$$

- Copper(II) oxide has *lost oxygen* and has been *reduced*.
- Hydrogen has *gained oxygen* and has been *oxidised*.

Redox reactions involve **red**uction and **ox**idation. If one process happens, so must the other – if something is reduced, something else must be oxidised.

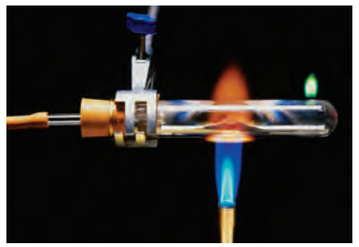
The terms oxidation and reduction are now applied to many reactions that do not involve oxygen. The modern definitions are in terms of either electrons or oxidation number.

Redox in terms of electrons

- Reduction is the gain of electrons.
- Oxidation is the loss of electrons.

The redox reaction below does not involve oxygen but does involve gain and loss of electrons.

$$2Fe(s) + 3Cl_2(g) \rightarrow 2FeCl_3(s)$$



▲ **Figure 1** *Reduction of copper(II) oxide with hydrogen – the green flame is excess hydrogen burning off*

Study tip

Remember **OILRIG**:

OIL **O**xidation **I**s **L**oss of electrons

RIG **R**eduction **I**s **G**ain of electrons.

$FeCl_3$ contains positive and negative ions, Fe^{3+} and Cl^-.

- Iron loses electrons and is oxidised $2Fe \rightarrow 2Fe^{3+} + 6e^-$
- Chlorine gains electrons and is reduced $3Cl_2 + 6e^- \rightarrow 6Cl^-$

The electrons gained and lost balance:

- 2 Fe in 2Fe each Fe loses $3e^-$ total of 6 electrons lost
- 6 Cl in $3Cl_2$ each Cl gains $1e^-$ total of 6 electrons gained

Redox in terms of oxidation number

- Reduction is a decrease in oxidation number.
- Oxidation is an increase in oxidation number.

The reaction below shows oxidation and reduction in terms of oxidation number. Each atom is assigned an oxidation number using the oxidation number rules:

$$Cu(s) + 2AgNO_3(aq) \rightarrow 2Ag(aq) + Cu(NO_3)_2(aq)$$

0 → +2 oxidation

+1 → 0 reduction

The changes in oxidation number apply to each atom and the total changes in oxidation number balance:

- 1 Cu in Cu Cu increases by +2 total increase = +2
- 2 Ag in $2AgNO_3$ each Ag decreases by −1 total decrease = −2

Redox reactions of acids

In Topic 4.1, you saw how acids produce salts in neutralisation reactions. Dilute acids also undergo redox reactions with some metals to produce salts and hydrogen gas.

metal + acid → salt + hydrogen

Reaction of zinc with dilute hydrochloric acid

$$Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$$

0 → +2 oxidation

+1 → 0 reduction

- 1 Zn in Zn Zn increases by +2 total increase = +2
- 2 H in 2HCl each H decreases by −1 total decrease = −2

The overall decrease in oxidation number of 2H = 2 × −1 = −2 balances the increase of Zn by +2.

Reaction of aluminium with dilute sulfuric acid

$$2Al(s) + 3H_2SO_4(aq) \rightarrow Al_2(SO_4)_3(aq) + 6H_2(g)$$

0 → +3 oxidation

+1 → 0 reduction

- 2 Al in 2Al each Al increases by +3 total increase = +6
- 6 H in $3H_2SO_4$ each H decreases by −1 total decrease = −6



▲ **Figure 2** *Iron reacts with chlorine gas in a gas jar, forming iron(III) chloride, $FeCl_3$*

Study tip

The oxidation number applies to each atom of an element. In this example, each atom of Ag changes from +1 to 0. The total change is −2 because there are 2 Ag, each changing by −1.

Summary questions

1 State the oxidation state of the species in the following:
- **a** Ag^+ *(1 mark)*
- **b** F_2 *(1 mark)*
- **c** $NaClO_3$ *(1 mark)*

2 State the oxidation number of sulfur in the following:
- **a** H_2S *(1 mark)*
- **b** SO_4^{2-} *(1 mark)*
- **c** $Na_2S_2O_3$ *(1 mark)*

3 The following reaction is a redox process:

$Mg + 2HCl \rightarrow MgCl_2 + H_2$

- **a** Identify the changes in oxidation number. *(2 marks)*
- **b** Explain which species is being oxidised and which is being reduced. *(2 marks)*

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1 25.0 cm^3 sample of an aqueous H_2SO_4 solution of unknown concentration is titrated with a 0.125 mol dm^{-3} NaOH. 22.40 cm^3 of NaOH are required to reach the end point.

$H_2SO_4 + 2NaOH \rightarrow Na_2SO_4 + 2H_2O$

a Calculate the amount, in mol of NaOH used. (*1 mark*)

b Calculate the amount, in mol of H_2SO_4 used. (*1 mark*)

c Calculate the concentration in mol dm^{-3} of the H_2SO_4. (*1 mark*)

2 The reaction below is a redox reaction.

$3CuO + 2NH_3 \rightarrow 3Cu + 3H_2O + N_2$

Explain in terms of oxidation number, what has been oxidised and what has been reduced. (*2 marks*)

3 This question looks at two redox reactions.

a $3Mg + 2Fe(NO_3)_3 \rightarrow 3Mg(NO_3)_2 + 2Fe$

(i) Explain in terms of electrons, what has been oxidised and what has been reduced. (*2 marks*)

(ii) What is the systematic name for $Fe(NO_3)_3$? (*1 mark*)

b $MnO_2 + 4HCl \rightarrow MnCl_2 + Cl_2 + 2H_2O$

(i) Explain, in terms of oxidation numbers, what has been reduced. (*2 marks*)

(ii) Use oxidation numbers to show that chlorine has only partly been reduced. (*2 marks*)

4 a A solution of calcium chloride can be prepared by neutralisation reactions of dilute hydrochloric acid with solid calcium carbonate, with solid calcium oxide and with aqueous calcium hydroxide.

(i) Write equations, with state symbols for these methods of preparing calcium chloride. (*6 marks*)

(ii) Why are these reactions all neutralisation reactions? (*1 mark*)

(iii) Write an ionic equation for the reaction of aqueous calcium hydroxide and hydrochloric acid. (*1 mark*)

b A solution of calcium chloride can be also prepared by the redox reaction of dilute hydrochloric acid and calcium metal.

(i) Explain reduction and oxidation in terms of electrons and oxidation numbers (*2 marks*)

(ii) Write an equation, with state symbols, for this reaction. (*2 marks*)

(iii) For this redox reaction, determine what had been oxidised and what has been reduced and identify the changes in oxidation numbers. (*2 marks*)

5 Tungsten ore contains WO_3. Tungsten can be extracted from WO_3 present in its ore in a redox reaction with hydrogen. The unbalanced equation is shown below.

$WO_3 + H_2 \rightarrow W + H_2O$

a Balance the equation. (*1 mark*)

b What is meant by oxidation and reduction in terms of electrons? (*1 mark*)

c Using oxidation numbers show that oxidation and reduction have taken place in this reaction. (*3 marks*)

d Some ore contains 2% of WO_3 by mass. Calculate the maximum mass of tungsten that could be obtained from the processing of 100 tonnes of ore. (*4 marks*)

6 A household cleaner containing dissolved ammonia is analysed as follows.

25.0 cm^3 of the cleaner is diluted to 250.0 cm^3. 25.0 cm^3 of the resulting solution is titrated with 0.125 mol dm^{-3} $H_2SO_4(aq)$ and 24.40 cm^3 were required to reach the end point. The equation is shown below.

$2NH_3(aq) + H_2SO_4(aq) \rightarrow (NH_4)_2SO_4(aq)$

a Calculate the amount, in mol, of H_2SO_4 used in the titration. (*1 mark*)

b Calculate the amount, in mol, of NH_3 used in the titration. (*1 mark*)

c Calculate the concentration of ammonia in the household cleaner

(i) in mol dm^{-3}; (*1 mark*)

(ii) in g dm^{-3}. (*1 mark*)

7 b The reaction between magnesium and sulfuric acid is a redox reaction.

$Mg(s) + H_2SO_4(aq) \rightarrow MgSO_4(aq) + H_2(g)$

(i) Use oxidation numbers to identify which element has been oxidised. Explain your answer. (*2 marks*)

(ii) Describe what you would see when magnesium reacts with an excess of sulfuric acid. (*2 marks*)

c Epsom salts can be used as bath salts to help relieve aches and pains. Epsom salts are crystals of hydrated magnesium sulfate, $MgSO_4 \bullet xH_2O$.

A sample of Epsom salts was heated to remove the water. 1.57 g of water was removed leaving behind 1.51 g of anhydrous $MgSO_4$.

(i) Calculate the amount, in mol, of anhydrous $MgSO_4$ formed. (*2 marks*)

(ii) Calculate the amount, in mol, of H_2O removed. (*1 mark*)

(iii) Calculate the value of x in $MgSO_4 \bullet xH_2O$. (*1 mark*)

F321 June 2009 1(b) (c)

8 A student carries out experiments using acids, bases and salts.

a Calcium nitrate, $Ca(NO_3)_2$, is an example of a salt.

The student prepares a solution of calcium nitrate by reacting dilute nitric acid, HNO_3, with the base calcium hydroxide, $Ca(OH)_2$.

(i) Why is calcium nitrate an example of a salt? (*1 mark*)

(ii) Write the equation for the reaction between dilute nitric acid and calcium hydroxide. Include state symbols. (*2 marks*)

(iii) Explain how the hydroxide ion in aqueous calcium hydroxide acts as a base when it neutralises dilute nitric acid. (*1 mark*)

b A student carries out a titration to find the concentration of some sulfuric acid.

The student finds that 25.00 cm^3 of 0.0880 mol dm^{-3} aqueous sodium hydroxide, NaOH, is neutralised by 17.60 cm^3 of dilute sulfuric acid, H_2SO_4.

$H_2SO_4(aq) + 2NaOH(aq) \rightarrow Na_2SO_4(aq) + 2H_2O(l)$

(i) Calculate the amount, in moles, of NaOH used. (*1 mark*)

(ii) Determine the amount, in moles, of H_2SO_4 used. (*1 mark*)

(iii) Calculate the concentration, in mol dm^{-3}, of the sulfuric acid. (*1 mark*)

c After carrying out the titration in (b), the student left the resulting solution to crystallise. White crystals were formed, with a formula of $Na_2SO_4 \bullet xH_2O$ and a molar mass of 322.1 g mol^{-1}.

(i) What term is given to the '$\bullet xH_2O$' part of the formula? (*1 mark*)

(ii) Using the molar mass of the crystals, calculate the value of x. (*2 marks*)

F321 Jan 10 Q2

9 Chemicals called 'acids' have been known throughout history. The word acid comes from the Latin 'acidus' meaning sour. Dilute sulfuric acid, H_2SO_4, is a common laboratory acid.

a (i) State the formulae of two ions released when sulfuric acid is in aqueous solution. (*2 marks*)

(ii) A student adds a sample of solid potassium carbonate, K_2CO_3, to an excess of dilute sulfuric acid.

Describe what the student would see and write the equation for the reaction which takes place. (*3 marks*)

b Dilute sulfuric acid reacts with alkalis such as sodium hydroxide. Solid sodium hydroxide is known as caustic soda. It has a household use as a drain cleaner. A student believes a box of caustic soda has been accidentally contaminated. To prove this, the student dissolves 2.00 g of the impure caustic soda in water and the solution is made up to 250 cm^3.

25.0 cm^3 of this solution of caustic soda is neutralised by 24.60 cm^3 of 0.100 mol dm^{-3} dilute sulfuric acid.

$H_2SO_4(aq) + 2NaOH(aq) \rightarrow Na_2SO_4(aq) + 2H_2O(l)$

(i) Calculate the amount, in moles, of H_2SO_4 used. (*1 mark*)

(ii) Determine the amount, in moles, of NaOH in the 25.0 cm^3 used. (*1 mark*)

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