PEARSON

# CHEMISTRY

2nd EDITION

## QUEENSLAND

## UNITS 3 & 4

Penny Commons
Chris Commons



## Skills and Assessment

QCE 2025 Chemistry SYLLABUS

# Contents

# Contents

# How to use this book

The *Pearson Chemistry Queensland Units 3 & 4 Skills and Assessment B*ook takes an intuitive, self-paced approach to science education that ensures every student has opportunities to practise, apply and extend their learning through a range of supportive and challenging activities.

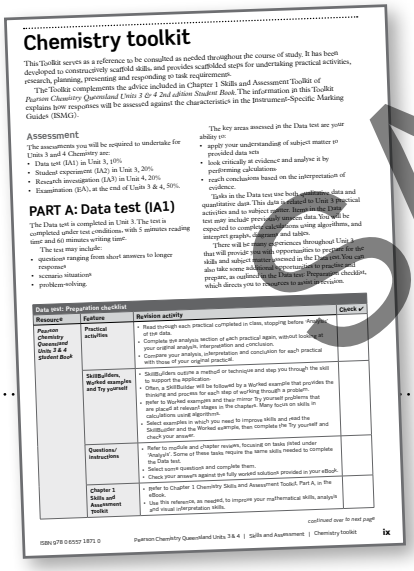
This resource has been developed by highly experienced and expert author teams, with leading Queensland specialists who have a working understanding of what teachers are looking for to support teaching and learning across the new Queensland Certificate of Education (QCE).

Fully written to the new QCE Unit 1–4 Syllabus, the Skills and Assessment Book is organised by units and topics, and the **unit opener** outlines the unit objectives that are addressed. The Skills and Assessment Book is further organised into topics. Each **topic** addresses all the subject matter and practical activities from the Syllabus.

All activities integrate into the *Pearson Chemistry Queensland Units 3 & 4 Student Book* for a complete teaching, learning and assessment program, making integration of practice and rich learning activities a seamless inclusion. The resource has been designed so that it can be used independently of the Student Book, providing flexibility in when and how the Skills and Assessment Book is engaged with.

## TOOLKIT

A complementary Toolkit supports the development of the skills and techniques needed to undertake Practical activities, the Data test, Student experiment and Research investigation, and also study skills. It also includes checklists and helpful hints to assist in fulfilling all assessment requirements.



Chemistry toolkit

This Toolkit serves as a reference to be consulted as needed throughout the course of study. It has been developed to constructively scaffold skills, and provides scaffolded steps for undertaking practical activities, research, planning, presenting and responding to task requirements.

The Toolkit complements the advice included in Chapter 1 Skills and Assessment Toolkit of *Pearson Chemistry Queensland Units 3 & 4 2nd edition Student Book*. The information in this Toolkit explains how responses will be assessed against the characteristics in the Instrument-Specific Marking Guides (ISMG).

Assessment

The assessments you will be required to undertake for Units 3 and 4 Chemistry are:

- Data test (IA1) in Unit 3, 10%
- Student experiment (IA2) in Unit 3, 20%
- Research investigation (IA3) in Unit 4, 20%
- Examination (EA), at the end of Units 3 & 4, 50%.

PART A: Data test (IA1)

The Data test is completed in Unit 3. The test is completed under test conditions, with 5 minutes reading time and 60 minutes writing time.

The test may include:

- questions ranging from short answers to longer responses
- scenario situations
- problem-solving.

The key areas assessed in the Data test are your ability to:

- apply your understanding of subject matter to provided data sets
- look critically at evidence and analyse it by performing calculations
- reach conclusions based on the interpretation of evidence.

Tasks in the Data test use both qualitative data and quantitative data. This data is related to Unit 3 practical activities and to subject matter from the Unit 3 content. The test may include previously unseen data. You will be expected to complete calculations using algorithms, and interpret graphs, diagrams and tables.

There will be many references throughout Unit 3 that will provide practice opportunities to prepare for the skills and subject matter assessed in the Data test. You can also take some additional opportunities to practise and prepare, as outlined in the Data test: Preparation checklist, which directs you to resources to assist participants.

Data test: Preparation checklist

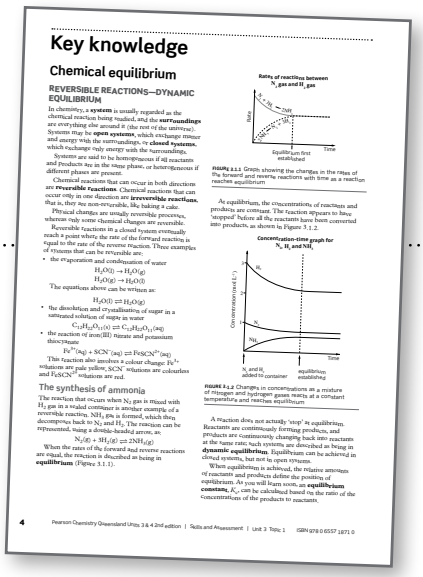
| Resource | Feature | Revision activity | Check ✓ |
|---|---|---|---|
| *Pearson Chemistry Queensland Units 3 & 4 Student Book* | Practical activities | • Read through each practical completed in class, stopping before Analysis of the data.<br>• Complete the analysis section of each practical again, without looking at your original analysis, interpretation and conclusion.<br>• Complete your analysis, interpretation and conclusion for each practical with those of your original practical. | |
| | SkillBuilders, Worked examples and Try yourself | • SkillBuilders outline a method or technique and step you through the skill to support the application.<br>• Often, a SkillBuilder will be followed by a Worked example that provides the thinking and process for each step of working through a problem.<br>• Refer to Worked examples and their mirror Try yourself problems that are placed at relevant stages in the chapter. Many focus on skills in calculations using algorithms.<br>• Select examples in which you need to improve skills and read the SkillBuilder and the Worked example, then complete the Try yourself and check your answer. | |
| | Questions/ instructions | • Refer to module and chapter reviews, focusing on tasks listed under 'Analysis'. Some of these tasks require the terms skills needed to complete the Data test.<br>• Select some questions and complete them.<br>• Check your answers against the fully worked solutions provided in your eBook. | |
| | Chapter 1 Skills and Assessment Toolkit | • Refer to Chapter 1 Chemistry Skills and Assessment Toolkit, Part A, in the eBook.<br>• Use this reference, as needed, to improve your mathematical skills, analysis and visual interpretation skills. | |

continued over to next page



## KEY KNOWLEDGE

Each topic begins with a key knowledge section. Key knowledge consists of a set of succinct summary notes that cover the subject matter for each topic of the Syllabus. This section is highly illustrative and written in a straightforward style to assist students of all abilities in focusing on the salient points. Key terms are bolded for ease of navigation and are reflected in the Student Book glossary. The key knowledge also serves as a ready reference when completing worksheets and practical activities and provides a handy set of revision and study notes.



Key knowledge

Chemical equilibrium

REVERSIBLE REACTIONS—DYNAMIC EQUILIBRIUM

In chemistry, a **system** is usually regarded as the chemical reaction being studied, and the **surroundings** are everything else around it (the rest of the universe). Systems may be **open systems**, which exchange matter and energy with the surroundings, or **closed systems**, which exchange only energy with the surroundings.

Systems are said to be homogeneous if all reactants and products are in the same phase, or heterogeneous if different phases are present.

Chemical reactions that can occur in both directions are **reversible reactions**. Chemical reactions that can occur only in one direction are **irreversible reactions**. That is, they are non-reversible, like baking a cake.

Physical changes are usually reversible processes, whereas only some chemical changes are reversible.

Reversible reactions in a closed system eventually reach a point where the rate of the forward reaction is equal to the rate of the reverse reaction. Three examples of systems that can be reversible are:

- the evaporation and condensation of water
  $H_2O(l) \rightarrow H_2O(g)$
  $H_2O(g) \rightarrow H_2O(l)$
  The equation above can be written as:
  $H_2O(l) \rightleftharpoons H_2O(g)$
- the dissolution and crystallisation of sugar in a saturated solution of sugar in water
  $C_{12}H_{22}O_{11}(s) \rightleftharpoons C_{12}H_{22}O_{11}(aq)$
- the reaction of iron(III) nitrate and potassium thiocyanate
  $Fe^{3+}(aq) + SCN^-(aq) \rightleftharpoons FeSCN^{2+}(aq)$
  This reaction also produces a colour change. $Fe^{3+}$ solutions are pale yellow, $SCN^-$ solutions are colourless and $FeSCN^{2+}$ solutions are red.

The synthesis of ammonia

The reaction that occurs when $N_2$ gas is mixed with $H_2$ gas in a sealed container is another example of a reversible reaction. $NH_3$ gas is formed, which then decomposes back to $N_2$ and $H_2$. The reaction can be represented using a double-headed arrow, as:

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

When the rates of the forward and reverse reactions are equal, the reaction is described as being in **equilibrium** (Figure 3.1.1).

Rate of reaction between $N_2$ gas and $H_2$ gas

Rate

Time

Equilibrium first established

FIGURE 3.1.1 Graph showing the changes in the rates of the forward and reverse reactions with time as a reaction reaches equilibrium

As equilibrium, the concentrations of reactants and products are constant. The reaction appears to have 'stopped' before all the reactants have been converted into products, as shown in Figure 3.1.2.

Concentration–time graph for $N_2$, $H_2$ and $NH_3$

Concentration (mol L$^{-1}$)

Time

$N_2$ and $H_2$ added to container

equilibrium establishing

FIGURE 3.1.2 Changes in concentrations as a mixture of nitrogen and hydrogen gases reacts at a constant temperature and reaches equilibrium

A reaction does not actually 'stop' at equilibrium. Reactants are continuously forming products, and products are continuously changing back into reactants at the same rate; such systems are described as being in **dynamic equilibrium**. Equilibrium can be achieved in closed systems, but not in open systems.

When equilibrium is achieved, the relative amounts of reactants and products define the position of equilibrium. As you will learn soon, an **equilibrium constant**, $K_c$, can be calculated based on the ratio of the concentrations of the products to reactants.

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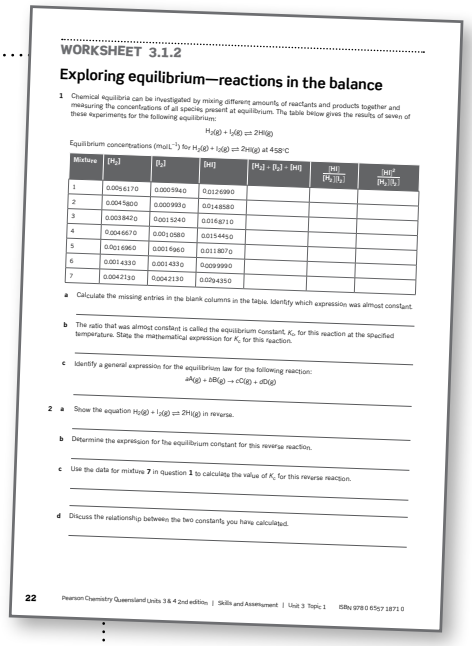
## WORKSHEETS

A diverse offering of instructive and self-contained worksheets is included in each topic. Common to all topics are the initial 'Knowledge preview' worksheets to activate prior knowledge; a 'Literacy review' worksheet to explicitly build language and application of scientific terminology; and finally, a 'Thinking about my learning' worksheet, which encourages students to reflect on their learning and identify areas for improvement. Other worksheets, with their range of activities and tasks, focus on the application of subject matter to assist in the consolidation of learning and the making of connections between subject matter.

Worksheets may be used as formative assessment and are clearly aligned to the Syllabus. A range of questions provide scaffolded support in alignment with the revised Syllabus. These worksheets build students' proficiency in skills and knowledge required to respond effectively to assessment requirements.



WORKSHEET 3.1.2

**Exploring equilibrium—reactions in the balance**

1 Chemical equilibria can be investigated by mixing different amounts of reactants and products together and measuring the concentrations of all species present at equilibrium. The table below gives the results of seven of these experiments for the following equilibrium:

$H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$

Equilibrium concentrations ($mol\,L^{-1}$) for $H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$ at 458°C

| Mixture | $[H_2]$ | $[I_2]$ | $[HI]$ | $[H_2] \times [I_2] \div [HI]$ | $\frac{[HI]}{[H_2][I_2]}$ | $\frac{[HI]^2}{[H_2][I_2]}$ |
|---|---|---|---|---|---|---|
| 1 | 0.005 6170 | 0.000 5940 | 0.012 6990 | | | |
| 2 | 0.004 5800 | 0.000 9930 | 0.014 8580 | | | |
| 3 | 0.003 8420 | 0.001 5240 | 0.016 8710 | | | |
| 4 | 0.004 6670 | 0.001 0580 | 0.015 4450 | | | |
| 5 | 0.001 6960 | 0.001 6960 | 0.011 8070 | | | |
| 6 | 0.001 4330 | 0.001 4330 | 0.009 9990 | | | |
| 7 | 0.004 2130 | 0.004 2130 | 0.029 4350 | | | |

a Calculate the missing entries in the blank columns in the table. Identify which expression was almost constant.

b The ratio that was almost constant is called the equilibrium constant, $K_c$, for this reaction at the specified temperature. State the mathematical expression for $K_c$ for this reaction.

c Identify a general expression for the equilibrium law for the following reaction:

$aA(g) + bB(g) \rightarrow cC(g) + dD(g)$

2 a Show the equation $H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$ in reverse.

b Determine the expression for the equilibrium constant for this reverse reaction.

c Use the data for mixture **7** in question **1** to calculate the value of $K_c$ for this reverse reaction.

d Discuss the relationship between the two constants you have calculated.

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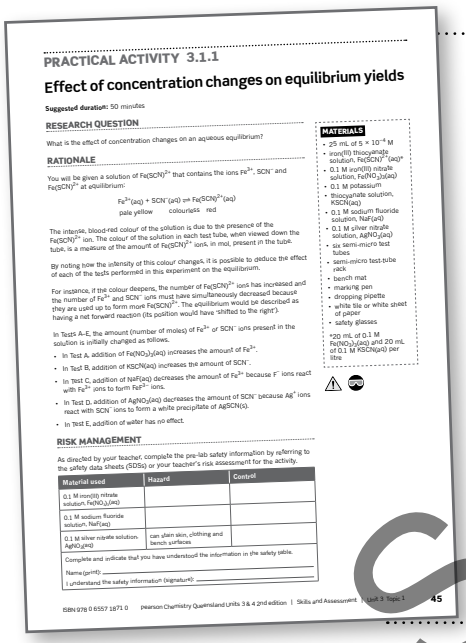

PRACTICAL ACTIVITY 3.1.1

**Effect of concentration changes on equilibrium yields**

**Suggested duration:** 50 minutes

**RESEARCH QUESTION**

What is the effect of concentration changes on an aqueous equilibrium?

**RATIONALE**

You will be given a solution of $Fe(SCN)^{2+}$ that contains the ions $Fe^{3+}$, $SCN^-$ and $Fe(SCN)^{2+}$ at equilibrium:

$Fe^{3+}(aq) + SCN^-(aq) \rightleftharpoons Fe(SCN)^{2+}(aq)$

pale yellow colourless red

The intense, blood-red colour of the solution is due to the presence of the $Fe(SCN)^{2+}$ ion. The colour of the solution in each test tube, when viewed down the tube, is a measure of the amount of $Fe(SCN)^{2+}$ ions in mol, present in the tube.

By noting how the intensity of this colour changes, it is possible to deduce the effect of each of the tests performed in this experiment on the equilibrium.

For instance, if the colour deepens, the number of $Fe(SCN)^{2+}$ ions has increased and the number of $Fe^{3+}$ and $SCN^-$ ions must have simultaneously decreased because they are used up to form more $Fe(SCN)^{2+}$. The equilibrium would be described as having a net forward reaction (its position would have 'shifted to the right').

In Tests A–E, the amount (number of moles) of $Fe^{3+}$ or $SCN^-$ ions present in the solution is initially changed as follows.

- In Test A, addition of $Fe(NO_3)_3(aq)$ increases the amount of $Fe^{3+}$.
- In Test B, addition of KSCN(aq) increases the amount of $SCN^-$.
- In Test C, addition of NaF(aq) decreases the amount of $Fe^{3+}$ because $F^-$ ions react with $Fe^{3+}$ ions to form $FeF^{2+}$ ions.
- In Test D, addition of $AgNO_3(aq)$ decreases the amount of $SCN^-$ because $Ag^+$ ions react with $SCN^-$ ions to form a white precipitate of AgSCN(s).
- In Test E, addition of water has no effect.

**MATERIALS**

- 25 mL of $5 \times 10^{-4}$ M iron(III) thiocyanate solution, $Fe(SCN)^{2+}(aq)$*
- 0.1 M iron(III) nitrate solution, $Fe(NO_3)_3(aq)$
- 0.1 M potassium thiocyanate solution, KSCN(aq)
- 0.1 M sodium fluoride solution, NaF(aq)
- 0.1 M silver nitrate solution, $AgNO_3(aq)$
- six semi-micro test tubes
- semi-micro test-tube rack
- bench mat
- marking pen
- dropping pipette
- white tile or white sheet of paper
- safety glasses

*20 mL of 0.1 M $Fe(NO_3)_3(aq)$ and 20 mL of 0.1 M KSCN(aq) per litre

**RISK MANAGEMENT**

As directed by your teacher, complete the pre-lab safety information by referring to the safety data sheets (SDSs) or your teacher's risk assessment for the activity.

| Material used | Hazard | Control |
|---|---|---|
| 0.1 M iron(III) nitrate solution, $Fe(NO_3)_3(aq)$ | | |
| 0.1 M sodium fluoride solution, NaF(aq) | | |
| 0.1 M silver nitrate solution, $AgNO_3(aq)$ | can stain skin, clothing and bench surfaces | |
| Complete and indicate that you have understood the information in the safety table. | | |
| Name (print): _____ | | |
| I understand the safety information (signature): _____ | | |

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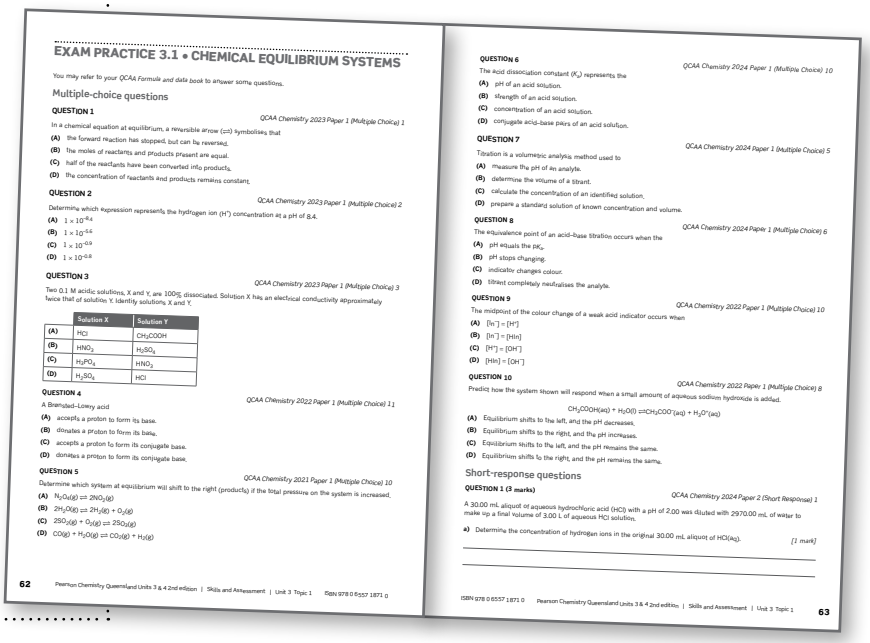
## PRACTICAL ACTIVITIES

Practical activities take a highly scaffolded approach from beginning to completion and give students the opportunity to complete practical work related to the subject matter covered in the Syllabus. Practical activities include a rich assortment of tasks that maximise the learning opportunities and build experience in skill application to perform calculations and analysis of data, which are necessary for the Data test. A diverse range of high-quality practical activities is included to provide depth and choice for teachers. Like the worksheets, a range of questions building from foundation to challenging is included, which are written to reflect the Syllabus cognitions at the expected level.



## EXAM PRACTICE QUESTIONS

Each topic concludes with a comprehensive set of question items consisting of multiple-choice and short-answer responses taken from past QCAA Chemistry examination papers. This provides students with exposure to, and the opportunity to practise, drawing together subject matter and skills to respond to examination-style assessment.



EXAM PRACTICE 3.1 • CHEMICAL EQUILIBRIUM SYSTEMS

You may refer to your *QCAA Formula and data book* to answer some questions.

**Multiple-choice questions**

**QUESTION 1** *QCAA Chemistry 2023 Paper 1 (Multiple Choice) 1*

In a chemical equation at equilibrium, a reversible arrow (⇌) symbolises that

**(A)** the forward reaction has stopped, but can be reversed.
**(B)** the moles of reactants and products present are equal.
**(C)** half of the reactants have been converted into products.
**(D)** the concentration of reactants and products remains constant.

**QUESTION 2** *QCAA Chemistry 2023 Paper 1 (Multiple Choice) 2*

Determine which expression represents the hydrogen ion ($H^+$) concentration at a pH of 8.4.

**(A)** $1 \times 10^{-8.4}$
**(B)** $1 \times 10^{-5.6}$
**(C)** $1 \times 10^{-0.9}$
**(D)** $1 \times 10^{0.8}$

**QUESTION 3** *QCAA Chemistry 2023 Paper 1 (Multiple Choice) 3*

Two 0.1 M acidic solutions, X and Y, are 100% dissociated. Solution X has an electrical conductivity approximately twice that of solution Y. Identify solutions X and Y.

| | Solution X | Solution Y |
|---|---|---|
| **(A)** | HCl | $CH_3COOH$ |
| **(B)** | $HNO_3$ | $H_2SO_4$ |
| **(C)** | $H_2PO_4$ | $HNO_3$ |
| **(D)** | $H_2SO_4$ | HCl |

**QUESTION 4** *QCAA Chemistry 2022 Paper 1 (Multiple Choice) 11*

A Brønsted–Lowry acid

**(A)** accepts a proton to form its base.
**(B)** donates a proton to form its base.
**(C)** accepts a proton to form its conjugate base.
**(D)** donates a proton to form its conjugate base.

**QUESTION 5** *QCAA Chemistry 2021 Paper 1 (Multiple Choice) 10*

Determine which system at equilibrium will shift to the right (products) if the total pressure on the system is increased.

**(A)** $N_2O_4(g) \rightleftharpoons 2NO_2(g)$
**(B)** $2H_2O(g) \rightleftharpoons 2H_2(g) + O_2(g)$
**(C)** $2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$
**(D)** $CO(g) + H_2O(g) \rightleftharpoons CO_2(g) + H_2(g)$

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**QUESTION 6** *QCAA Chemistry 2024 Paper 1 (Multiple Choice) 10*

The acid dissociation constant ($K_a$) represents the

**(A)** pH of an acid solution.
**(B)** strength of an acid solution.
**(C)** concentration of an acid solution.
**(D)** conjugate acid–base pairs of an acid solution.

**QUESTION 7** *QCAA Chemistry 2024 Paper 1 (Multiple Choice) 5*

Titration is a volumetric analysis method used to

**(A)** measure the pH of an analyte.
**(B)** determine the volume of a titrant.
**(C)** calculate the concentration of an identified solution.
**(D)** prepare a standard solution of known concentration and volume.

**QUESTION 8** *QCAA Chemistry 2024 Paper 1 (Multiple Choice) 6*

The equivalence point of an acid–base titration occurs when the

**(A)** pH equals the $pK_a$.
**(B)** pH stops changing.
**(C)** indicator changes colour.
**(D)** titrant completely neutralises the analyte.

**QUESTION 9** *QCAA Chemistry 2022 Paper 1 (Multiple Choice) 10*

The midpoint of the colour change of a weak acid indicator occurs when

**(A)** $[In^-] = [H^+]$
**(B)** $[In^-] = [HIn]$
**(C)** $[H^+] = [OH^-]$
**(D)** $[HIn] = [OH^-]$

**QUESTION 10** *QCAA Chemistry 2022 Paper 1 (Multiple Choice) 8*

Predict how the system shown will respond when a small amount of aqueous sodium hydroxide is added.

$CH_3COOH(aq) + H_2O(l) \rightleftharpoons CH_3COO^-(aq) + H_3O^+(aq)$

**(A)** Equilibrium shifts to the left, and the pH decreases.
**(B)** Equilibrium shifts to the right, and the pH increases.
**(C)** Equilibrium shifts to the left, and the pH remains the same.
**(D)** Equilibrium shifts to the right, and the pH remains the same.

**Short-response questions**

**QUESTION 1 (3 marks)** *QCAA Chemistry 2024 Paper 2 (Short Response) 1*

A 30.00 mL aliquot of aqueous hydrochloric acid (HCl) with a pH of 2.00 was diluted with 2970.00 mL of water to make up a final volume of 3.00 L of aqueous HCl solution.

a) Determine the concentration of hydrogen ions in the original 30.00 mL aliquot of $HCl(aq)$. *[1 mark]*

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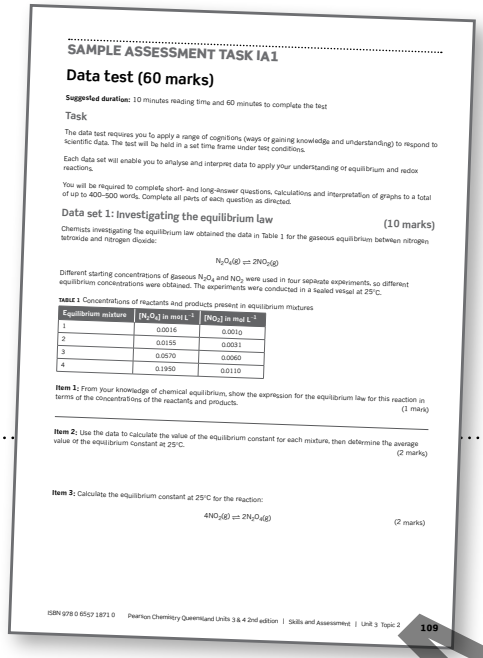
## SAMPLE ASSESSMENT TASKS

Sample assessment tasks for the Data test, Student experiment and Research investigation provide opportunities for students to practise responding to these assessment tasks. The activities are designed to support students by guiding and scaffolding them through each aspect of these assessments.



SAMPLE ASSESSMENT TASK IA1

**Data test (60 marks)**

**Suggested duration:** 10 minutes reading time and 60 minutes to complete the test

**Task**

The data test requires you to apply a range of cognitions (ways of gaining knowledge and understanding) to respond to scientific data. The test will be held in a set time frame under test conditions.

Each data set will enable you to analyse and interpret data to apply your understanding of equilibrium and redox reactions.

You will be required to complete short- and long-answer questions, calculations and interpretation of graphs to a total of up to 400–500 words. Complete all parts of each question as directed.

**Data set 1: Investigating the equilibrium law** **(10 marks)**

Chemists investigating the equilibrium law obtained the data in Table 1 for the gaseous equilibrium between nitrogen tetroxide and nitrogen dioxide:

$$N_2O_4(g) \rightleftharpoons 2NO_2(g)$$

Different starting concentrations of gaseous $N_2O_4$ and $NO_2$ were used in four separate experiments, so different equilibrium concentrations were obtained. The experiments were conducted in a sealed vessel at 25°C.

**TABLE 1** Concentrations of reactants and products present in equilibrium mixtures

| Equilibrium mixture | $[N_2O_4]$ in mol $L^{-1}$ | $[NO_2]$ in mol $L^{-1}$ |
|---|---|---|
| 1 | 0.0016 | 0.0010 |
| 2 | 0.0155 | 0.0031 |
| 3 | 0.0570 | 0.0060 |
| 4 | 0.1950 | 0.0110 |

**Item 1:** From your knowledge of chemical equilibrium, show the expression for the equilibrium law for this reaction in terms of the concentrations of the reactants and products. (1 mark)

**Item 2:** Use the data to calculate the value of the equilibrium constant for each mixture, then determine the average value of the equilibrium constant at 25°C. (2 marks)

**Item 3:** Calculate the equilibrium constant at 25°C for the reaction:

$$4NO_2(g) \rightleftharpoons 2N_2O_4(g)$$

(2 marks)

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## ICONS AND FEATURES

Many of the Practical activities are further supported by a complementary SPARKlab alternative practical.

Use the activities for practice, application and revision of subject matter.

The **safety icon** highlights significant hazards, indicating caution is needed.

The **safety glasses icon** highlights that protective eyewear is to be worn during the practical activity.

## RATE MY LEARNING

This innovative feature appears at the end of most of the worksheets, all practical activities and sample assessment tasks. It provides students with the opportunity for self-reflection and self-assessment. Students are encouraged to consider how they can continue to improve, and to identify areas of focus for further skill and subject matter development. This tool has been based on the syllabus requirements to provide teachers with a simple yet effective tool for gauging student needs.

| RATE MY LEARNING | | | | |
|---|---|---|---|---|
| • I get it.<br>• I can apply/teach it. | • I get it.<br>• I can show I get it. | • I almost get it.<br>• I might need help. | • I get some of it.<br>• I need help. | • I don't get it.<br>• I need lots of help. |

## Teacher support

Fully worked solutions, suggested answers and responses to sample assessment tasks, as well as practical activity support including full risk assessments, expected results and handy hints, are provided for teachers through the teacher support subscription.

# Chemistry toolkit

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| Resource | Feature | Revision activity | Check ✔ |
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| | **SkillBuilders, Worked examples and Try yourself** | • SkillBuilders outline a method or technique and step you through the skill to support the application.<br>• Often, a SkillBuilder will be followed by a Worked example that provides the thinking and process for each step of working through a problem.<br>• Refer to Worked examples and their mirror Try yourself problems that are placed at relevant stages in the chapters. Many focus on skills in calculations using algorithms.<br>• Select examples in which you need to improve skills and read the SkillBuilder and the Worked example, then complete the Try yourself and check your answer. | |
| | **Questions/ instructions** | • Refer to module and chapter reviews, focusing on tasks listed under 'Analysis'. Some of these tasks require the same skills needed to complete the Data test.<br>• Select some questions and complete them.<br>• Check your answers against the fully worked solutions provided in your eBook. | |
| | **Chapter 1 Skills and Assessment Toolkit** | • Refer to Chapter 1 Chemistry Skills and Assessment Toolkit, Part A, in the eBook.<br>• Use this reference, as needed, to improve your mathematical skills, analysis and visual interpretation skills. | |

*continued over to next page*

| Data test: Preparation checklist | | | |
|---|---|---|---|
| **Resource** | **Feature** | **Revision activity** | **Check ✔** |
| ***Pearson Chemistry Queensland Units 3 & 4 Skills and Assessment Book*** | **Practical activities** | • See suggestions and support for practicals; these include working with data. | |
| | **Exam review** | • Refer to the Exam review tasks for examples of the types of items on the Data test. | |
| | **Practice data test** | • Complete the practice Sample Assessment Task IA1: Data test provided in this book.<br>• Complete these tasks under test conditions. | |

# PART B: Student experiment (IA2)

The Student experiment uses practical investigation methodology including a research question or hypothesis developed by the student, the collection of primary data, and the analysis and synthesis of that data.

The Research question developed must:

- relate back to a practical related to the subject matter.

## CONDUCTING THE STUDENT EXPERIMENT

A great deal of preparation is needed before starting the experiment. Use the Student Experiment Checklist as a guide.

Refer to *Pearson Chemistry Queensland Units 3 & 4 Student Book*, Chapter 1 ebook.

- Part A covers Scientific skills such as mathematical basics, representations in graphics, tables and graphs, data analysis.
- Part B covers all aspects of the student experiment and includes a sample student report.



| Student Experiment Checklist | | | | |
|---|---|---|---|---|
| **Task** | | **Activity** | **Date due** | **Complete ✔** |
| **Form ideas and develop the experiment question** | **Initial practical** | • Identify the practical to be modified for your student's experiment. | | |
| | **Background/ Rationale** | • Commence a journal to record all aspects of the assessment task.<br>• Research relevant background information.<br>• By referring to both the original class experiment and to Unit 3 theory, describe your reason for conducting the experiment.<br>Refer to sources and secondary data | | |
| | **Variables** | • Understand the data of the original experiment.<br>• Identify the independent and dependent variables of the class practical.<br>• Understand the relationship between the dependent and independent variables shown in the findings of this investigation.<br>• identify the controlled variables and consider how they will be kept constant. | | |
| | **Modification** | • Modify the original experiment.<br>• Develop a new experiment research question to be investigated in your student experiment.<br>• Identify the dependent and independent variables and check whether they can be measured to provide data.<br>• Use your understanding of chemical concepts to identify change(s) to be made to the class practical—modify, refine, extend or redirect. | | |
| | **Justification** | • Justify/provide rationale for the modification of the new experiment question. | | |

| Student Experiment Checklist | | | | |
|---|---|---|---|---|
| **Task** | | **Activity** | **Date due** | **Complete ✔** |
| **Find the methodology and data** | **Methodology** | • Plan the method to be followed for the experiment.<br>• Write a statement of the aim that covers what the scientist wants to show, verify or find out in the experiment.<br>  • May be expressed as a statement or question.<br>  • May be one or two sentences.<br>  • Often written as 'To investigate the effect of ... on ...' or 'To investigate if a correlation exists between ... and ...'. | | |
| | **Risk assessment** | • Identify and manage risks and potential dangers by completing the Risk Assessment Form.<br>• Also consider ethical and environmental issues. | | |
| | **Materials** | • Make a list of all equipment, chemicals and materials used.<br>• Include quantities (chemicals) and sizes (equipment). | | |
| | **The experiment** | • Conduct the experiment.<br>• Collect the relevant data to answer the research question. | | |
| | **Results** | • Record all measurements taken during the experiment as well as observations.<br>• Collect sufficient and relevant data.<br>• Observations may be recorded as text, diagrams, photos or videos.<br>• Process data and present it correctly.<br>• Most common records of primary data are tables with titles and units.<br>• Most common records of processed data are tables, graphs and can include calculations.<br>• Check for errors and mistakes in the data collected. Take steps to reduce these. | | |
| **Analyse the evidence** | **Organise data collected** | • Process the data using mathematical techniques and graphs.<br>• Identify the trends, patterns or relationships.<br>• Identify the uncertainties and limitations of the evidence.<br>• Include the interpretation of results. | | |
| **Interpret and evaluate the evidence** | **Work with data collected; relate back to the experiment question** | • Draw conclusion/s from the evidence that address the experiment question.<br>• Evaluate the reliability and validity of the experimental process.<br>• Evaluate the method used by discussing reliability and validity.<br>• Comment on whether the results relate to the experiment question.<br>• Provide suggestions for improvements and extensions to the experiment. | | |
| **The experiment report** | **Presentation format** | • Decide on the presentation format: written or multi-modal.<br>• Check length requirements for your selected format. | | |
| | **Communication** | • Communicate ideas in scientific language and representations.<br>• Include in-text citations and reference list.<br>• Write using your own words to avoid plagiarising.<br>• Ensure length requirements are not exceeded. | | |

# UNIT 3 Equilibrium, acids and redox reactions

**TOPIC 1** Chemical equilibrium systems

**TOPIC 2** Oxidation and reduction

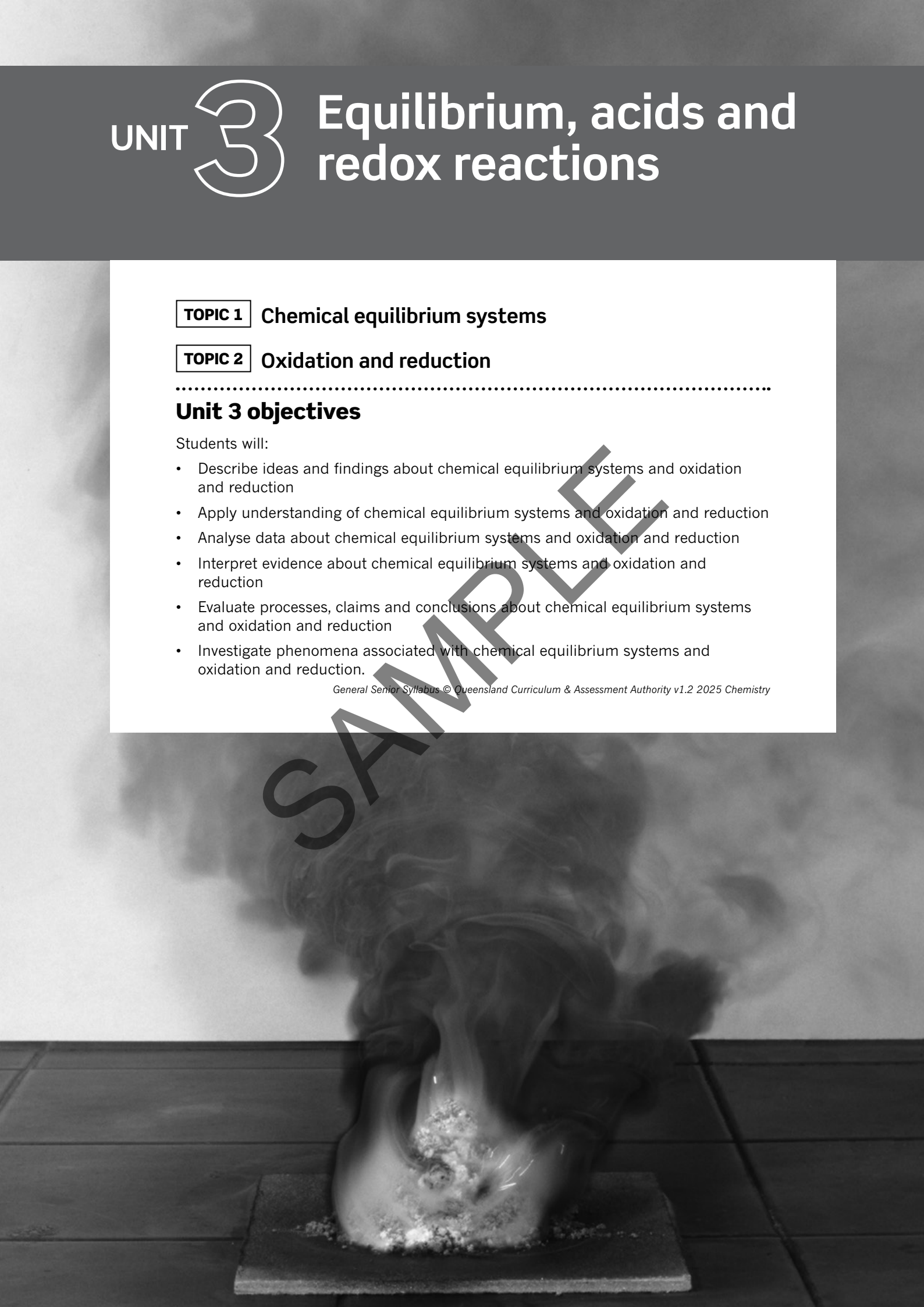
## Unit 3 objectives

Students will:

- Describe ideas and findings about chemical equilibrium systems and oxidation and reduction
- Apply understanding of chemical equilibrium systems and oxidation and reduction
- Analyse data about chemical equilibrium systems and oxidation and reduction
- Interpret evidence about chemical equilibrium systems and oxidation and reduction
- Evaluate processes, claims and conclusions about chemical equilibrium systems and oxidation and reduction
- Investigate phenomena associated with chemical equilibrium systems and oxidation and reduction.







TOPIC

# 1 Chemical equilibrium systems

- ☐ Worksheet 3.1.1 Knowledge preview—thinking about rates and energy

## CHEMICAL EQUILIBRIUM

- ☐ Worksheet 3.1.2 Exploring equilibrium—reactions in the balance

## FACTORS THAT AFFECT EQUILIBRIUM

- ☐ Worksheet 3.1.3 Equilibrium—Le Châtelier's principle and the equilibrium law
- ☐ Practical activity 3.1.1 Effect of concentration changes on equilibrium yields
- ☐ Practical activity 3.1.2 Effect of temperature on equilibrium yields

## EQUILIBRIUM CONSTANTS

- ☐ Worksheet 3.1.4 Calculations—equilibrium constants and concentrations

## PROPERTIES OF ACIDS AND BASES

- ☐ Practical activity 3.1.3 Comparing the strengths of acids and bases
- ☐ Practical activity 3.1.4 Electrical conductivity of strong and weak acids

## pH SCALE

- ☐ Worksheet 3.1.5 Acid–base relationships—pH and pOH, p$K_a$ and p$K_b$
- ☐ Practical activity 3.1.3 Comparing the strengths of acids and bases
- ☐ Practical activity 3.1.5 Amphiprotic substances in water

## BRØNSTED–LOWRY MODEL

- ☐ Practical activity 3.1.5 Amphiprotic substances in water

## DISSOCIATION CONSTANTS

- ☐ Worksheet 3.1.5 Acid–base relationships—pH and pOH, p$K_a$ and p$K_b$
- ☐ Worksheet 3.1.7 Acid–base and conductometric curves—detecting an equivalence point

## ACID–BASE INDICATORS

- ☐ Worksheet 3.1.6 Indicators and $K_a$—colourful considerations

## VOLUMETRIC ANALYSIS

- ☐ Worksheet 3.1.7 Acid–base and conductometric curves—detecting an equivalence point
- ☐ Worksheet 3.1.8 Volumetric analysis—finding the fizz in soda water
- ☐ Practical activity 3.1.6 Determination of the ethanoic acid concentration of vinegar



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## BUFFERS

- ☐ Worksheet 3.1.9 Buffers—keeping pH in control
- ☐ Worksheet 3.1.10 Literacy review—equilibrium terms and expressions
- ☐ Worksheet 3.1.11 Thinking about my learning

**Exam practice 3.1**

# Key knowledge

## Chemical equilibrium

### REVERSIBLE REACTIONS—DYNAMIC EQUILIBRIUM

In chemistry, a **system** is usually regarded as the chemical reaction being studied, and the **surroundings** are everything else around it (the rest of the universe). Systems may be **open systems**, which exchange matter and energy with the surroundings, or **closed systems**, which exchange only energy with the surroundings.

Systems are said to be homogeneous if all reactants and products are in the same phase, or heterogeneous if different phases are present.

Chemical reactions that can occur in both directions are **reversible reactions**. Chemical reactions that can occur only in one direction are **irreversible reactions**, that is, they are non-reversible, like baking a cake.

Physical changes are usually reversible processes, whereas only some chemical changes are reversible.

Reversible reactions in a closed system eventually reach a point where the rate of the forward reaction is equal to the rate of the reverse reaction. Three examples of systems that can be reversible are:

- the evaporation and condensation of water

$$H_2O(l) \rightarrow H_2O(g)$$
$$H_2O(g) \rightarrow H_2O(l)$$

The equations above can be written as:

$$H_2O(l) \rightleftharpoons H_2O(g)$$

- the dissolution and crystallisation of sugar in a saturated solution of sugar in water

$$C_{12}H_{22}O_{11}(s) \rightleftharpoons C_{12}H_{22}O_{11}(aq)$$

- the reaction of iron(III) nitrate and potassium thiocyanate

$$Fe^{3+}(aq) + SCN^-(aq) \rightleftharpoons FeSCN^{2+}(aq)$$

This reaction also involves a colour change: $Fe^{3+}$ solutions are pale yellow, $SCN^-$ solutions are colourless and $FeSCN^{2+}$ solutions are red.

### The synthesis of ammonia

The reaction that occurs when $N_2$ gas is mixed with $H_2$ gas in a sealed container is another example of a reversible reaction. $NH_3$ gas is formed, which then decomposes back to $N_2$ and $H_2$. The reaction can be represented, using a double-headed arrow, as:

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

When the rates of the forward and reverse reactions are equal, the reaction is described as being in **equilibrium** (Figure 3.1.1).

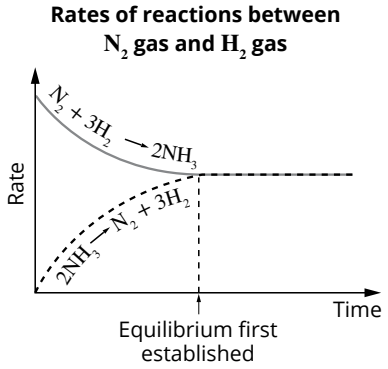


**FIGURE 3.1.1** Graph showing the changes in the rates of the forward and reverse reactions with time as a reaction reaches equilibrium

At equilibrium, the concentrations of reactants and products are constant. The reaction appears to have 'stopped' before all the reactants have been converted into products, as shown in Figure 3.1.2.

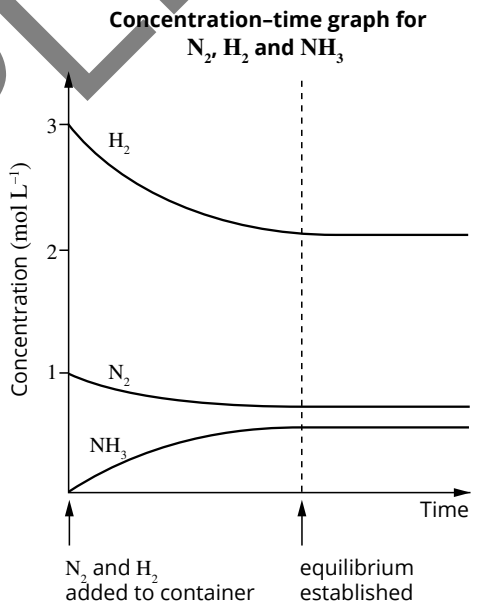


**FIGURE 3.1.2** Changes in concentrations as a mixture of nitrogen and hydrogen gases reacts at a constant temperature and reaches equilibrium

A reaction does not actually 'stop' at equilibrium. Reactants are continuously forming products, and products are continuously changing back into reactants at the same rate; such systems are described as being in **dynamic equilibrium**. Equilibrium can be achieved in closed systems, but not in open systems.

When equilibrium is achieved, the relative amounts of reactants and products define the position of equilibrium. As you will learn soon, an **equilibrium constant**, $K_c$, can be calculated based on the ratio of the concentrations of the products to reactants.



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## Energy profile diagrams and reversible reactions

When reactions occur, bonds of reactant particles are broken and the new bonds of the product particles are formed. Energy profile diagrams, such as Figure 3.1.3, show the energy changes that occur during a reaction. These diagrams allow you to understand why reactions can be reversible. When the particles collide with sufficient energy to break bonds in the reacting particles, the products form. The reverse process can also occur, in which product particles absorb energy to break their bonds and re-form the reactants.

In Figure 3.1.3, the forward reaction is **exothermic** (energy-releasing) and the reverse reaction is **endothermic** (energy-absorbing). The energy that is required for the reaction to occur is called the **activation energy** and is represented by the peak of the energy profile diagram. The activation energy for the endothermic reaction is the sum of the $\Delta H$ for the reaction and the activation energy for the exothermic reaction.

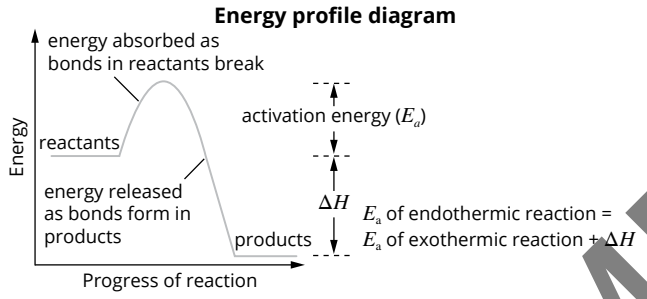


**FIGURE 3.1.3** Energy profile diagram for an exothermic reaction

# Factors that affect equilibrium

## LE CHÂTELIER'S PRINCIPLE

**Le Châtelier's principle** states: 'If a change is imposed on a system at equilibrium, the system will adjust itself to partially oppose the effect of the change'.

The effect of a change on an equilibrium can be predicted by Le Châtelier's principle. You can understand the basis of Le Châtelier's principle using **collision theory**: a reaction can occur when reactant particles collide with sufficient energy and the appropriate molecular orientation. When the reactants are mixed initially and the reaction commences, the number and concentration of product particles will increase. However, as the frequency of collisions between the product particles increases, the rate of reformation of reactant particles also increases. Eventually, the rates of the forward and reverse reactions become equal, and equilibrium is established.

We will now explore the effect of different changes on equilibrium systems.

## Effect of concentration change on equilibrium

Consider a container of $N_2$, $H_2$ and $NH_3$ at equilibrium:

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

If extra nitrogen gas is added to the container, without any change in the volume or temperature, the mixture will momentarily not be in equilibrium. The system will adjust to form a new equilibrium with different concentrations of $N_2$, $H_2$ and $NH_3$.

Figure 3.1.4(a) on page 6 is a concentration–time graph showing the result of adding $N_2$ to the initial equilibrium. As predicted by Le Châtelier's principle, you can see that adding more reactant to the equilibrium results in the production of more $NH_3$ as the equilibrium position shifts to use up some of the added $N_2$.

The rate–time graph in Figure 3.1.4(b) on page 6 shows the effects on the rate of the forward and reverse reactions when the extra nitrogen is added. Initially, there is a net forward reaction due to the added nitrogen. Then the rates of the forward and reverse reactions become equal again. Overall, the equilibrium position has shifted towards the right, and more $NH_3$ has been formed.

## Effect of overall pressure change on equilibrium

When the overall pressure of an equilibrium system involving gases is decreased by increasing the volume, the system restores equilibrium by increasing the overall pressure. This is done by moving in the direction of the most particles. In the example:

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

a decrease in pressure causes the system to move in the direction of the reactants (a net reverse reaction), because there are 4 moles of gas on the reactant side and only 2 moles of gas on the product side.

## Effect of dilution on equilibrium

Now consider the aqueous equilibrium system:

$$Fe^{3+}(aq) + SCN^-(aq) \rightleftharpoons FeSCN^{2+}(aq)$$

Figure 3.1.5 on page 6 shows the effect of doubling the volume of the system by adding water without any change in temperature. Dilution causes the concentration of ions in solution to decrease. Le Châtelier's principle tells us that equilibrium will be re-established by the system by shifting in the direction that increases the ion concentration. In this example, there are 2 moles of ions on the reactant side and 1 mole of ions on the product side, so the position of equilibrium shifts in the direction of the reactants and there is a net reverse reaction.

Note that, although the equilibrium position shifts to the left, the concentrations of $Fe^{3+}(aq)$ and $SCN^-(aq)$ at the new equilibrium are lower than their concentrations prior to dilution. This is because the shift in the equilibrium only partially opposes the change.

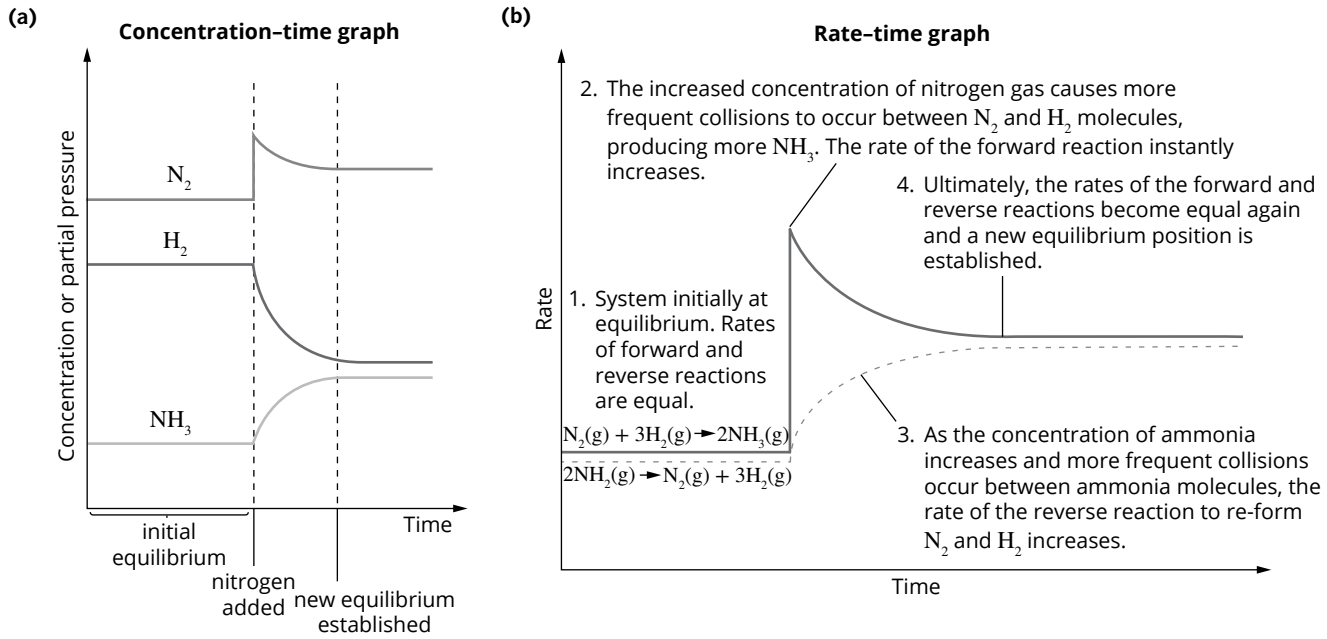


**FIGURE 3.1.4** (a) A representation of changes in concentrations that occur when additional nitrogen gas is added. (b) The effects on the rate of the forward and reverse reactions when more nitrogen is added.

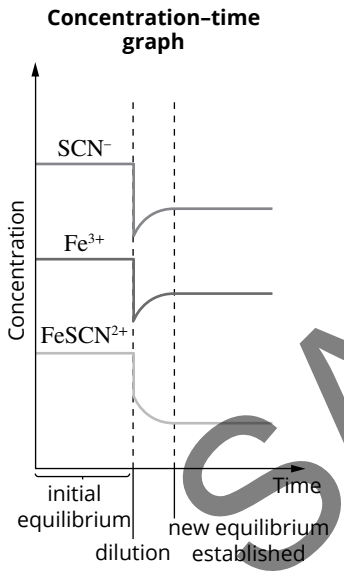


**FIGURE 3.1.5** Effect of dilution on the equilibrium: $Fe^{3+}(aq) + SCN^{-}(aq) \rightleftharpoons FeSCN^{2+}(aq)$

## Effect of a catalyst on equilibrium

Adding a catalyst to an equilibrium system speeds up the rates of the forward and backward reactions equally. As a result, the catalyst does not change the equilibrium concentrations of the reactants or products.

Table 3.1.1 summarises the effect of various changes on an equilibrium system when the temperature is kept constant in terms of Le Châtelier's principle and collision theory.

## Effect of temperature change on equilibrium

A rise in temperature decreases the amount of product at equilibrium in an exothermic reaction and increases the amount of product in an endothermic reaction (Table 3.1.2).



**TABLE 3.1.1** The effect of changes on an equilibrium system when the temperature is constant

| Change to system in equilibrium | Effect on equilibrium position predicted by Le Châtelier's principle | Collision theory explanation of the effect of change on equilibrium position |
|---|---|---|
| add extra reactant | shifts to the right (net forward reaction) | Collisions between reactant molecules are more frequent. Initially, the rate of the forward reaction increases. Then there is a gradual increase in the rate of the reverse reaction and a gradual decrease in the rate of the forward reaction until equilibrium is re-established. |
| add extra product | shifts to the left (net reverse reaction) | The rate of the reverse reaction increases initially. Then there is a gradual increase in the rate of the forward reaction and a gradual decrease in the rate of the reverse reaction until equilibrium is re-established. |
| decrease the pressure by increasing volume (for gases) | shifts in the direction of the most particles | Collisions are less frequent overall. The reaction that is less dependent on collisions (fewer particles reacting) initially occurs to a greater extent before the rates of the forward and reverse reactions become equal again. |

| Change to system in equilibrium | Effect on equilibrium position predicted by Le Châtelier's principle | Collision theory explanation of the effect of change on equilibrium position |
|---|---|---|
| add a catalyst | no change | No change in either direction, because the rates of the forward and reverse reactions are increased equally. The frequencies of successful collisions are equal. |
| add an inert gas, e.g. neon (container volume remains constant) | no change | There is no change in the concentration of reactant or product molecules, so no change in the frequency of collisions between reactants or products or rate of reaction. |
| add water (dilution of solutions) | shifts in the direction of the most particles | Collisions are less frequent overall. The reaction that is less dependent on collisions (fewer particles reacting) initially occurs to a greater extent before the rates become equal again. |

**TABLE 3.1.2** Relationship between concentration of products and the enthalpy change of a reaction, $\Delta H$, as temperature increases

| Sign of ΔH | Effect on concentration of products as temperature increases | Effect on $K_c$ as temperature increases |
|---|---|---|
| – (exothermic) | decreases | decreases |
| + (endothermic) | increases | increases |

The effect of a change in temperature on an equilibrium system is described in Table 3.1.3 in terms of Le Châtelier's principle and collision theory.

**TABLE 3.1.3** The effect of a change on an equilibrium system when the temperature is changed

| Change to system in equilibrium | Effect on equilibrium position predicted by Le Châtelier's principle | Collision theory explanation of the effect of change on equilibrium position |
|---|---|---|
| increasing the temperature for exothermic reactions | shifts to the left (to absorb some of the added energy) | All molecules have more energy and move faster. The direction of the endothermic reaction is favoured because it requires more energy to occur. Hence, the reverse reaction is favoured. |
| increasing the temperature for endothermic reactions | shifts to the right (to absorb some of the added energy) | All molecules have more energy and move faster. The direction of the endothermic reaction is favoured because it requires more energy to occur. Hence, the forward reaction is favoured. |

When the temperature increases, the rates of the forward and reverse reactions of an equilibrium system increase. One of these reactions will be endothermic, and the other will be exothermic. The endothermic reaction has a greater activation energy and is favoured by an increase in temperature, because a larger proportion of molecules will have sufficient energy to overcome the activation energy (Figure 3.1.6). The rate of the endothermic reaction therefore increases more than the rate of the exothermic reaction, and the position of equilibrium moves in the direction of the endothermic reaction. The reverse is true for exothermic reactions.

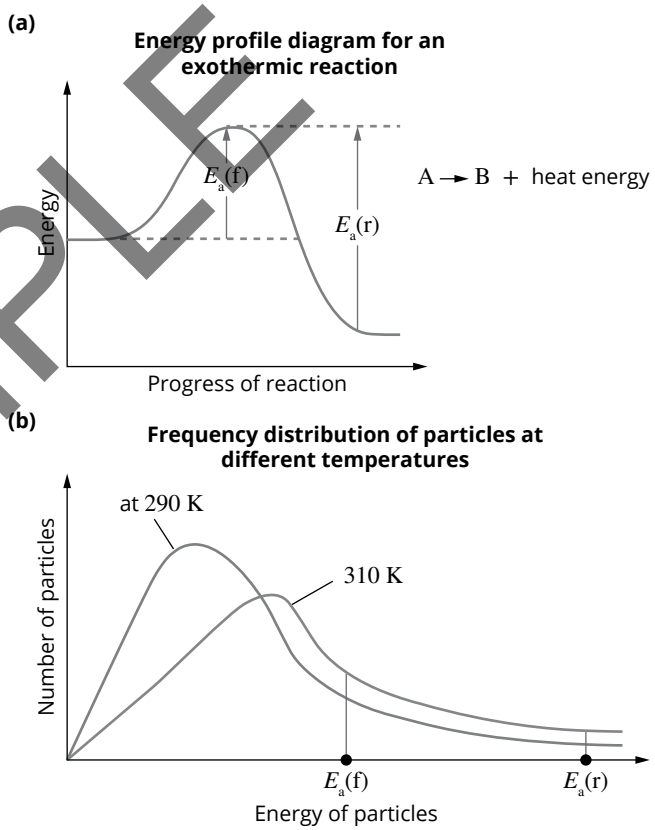


**FIGURE 3.1.6** (a) Energy profile diagram for an exothermic reaction. The activation energy of the forward reaction, $E_a$(f), is less than the activation energy of the reverse reaction, $E_a$(r). (b) The frequency distribution of molecules at two different temperatures. At a higher temperature, a greater proportion of particles have the necessary activation energy for the endothermic reverse reaction.

Consider the exothermic system when brown $NO_2$ gas is converted to colourless $N_2O_4$ gas:

$$2NO_2(g) \rightleftharpoons N_2O_4(g) \qquad \Delta H = -58\,kJ\,mol^{-1}$$

An increase in temperature causes a net reverse reaction to occur. For an equilibrium system in a sealed container, the concentration of $NO_2$ will increase and the concentration of $N_2O_4$ will decrease, as shown in Figure 3.1.7 on page 8.

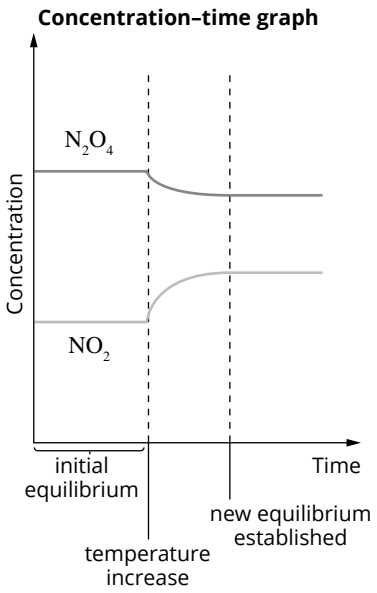


**FIGURE 3.1.7** The effect of a temperature increase on the equilibrium, $2NO_2(g) \rightleftharpoons N_2O_4(g)$, in a sealed container

# Equilibrium constants

At equilibrium, the rate of the forward reaction is equal to the rate of the reverse reaction. The point at which equilibrium is reached is not the same for all reactions. Some reactions go almost to completion, so that when equilibrium is reached, almost all the reactants have been converted into products. Other reactions yield almost no products, and most of the reactants are present at any instant. The term **position of equilibrium** is used to describe the extent to which the reaction has proceeded, or the **extent of reaction**.

## THE EQUILIBRIUM LAW

The relative concentrations of reactants and products at equilibrium can be described using the **equilibrium law**. At a specified temperature, the equilibrium law for the general reaction:

$$aA + bB \rightleftharpoons cC + dD$$

is given by:

$$K_c = \frac{[C]^c[D]^d}{[A]^a[B]^b}$$

The equilibrium constant, $K_c$, is the ratio of the equilibrium concentrations of the products to the equilibrium concentrations of the reactants, raised to powers that are their coefficients in the equation. The ratios of reactants to products when equilibrium is reached are different for different reactions.

The ratio of products to reactants can be calculated at any time during the reaction and is called the **reaction quotient**, $Q$, or concentration fraction.

- If $Q > K_c$, the reaction shifts left to establish equilibrium.
- If $Q < K_c$, the reaction shifts right to establish equilibrium.
- If $Q = K_c$, the reaction is at equilibrium.

For example, at equilibrium for the reaction:

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

$$Q = K_c = \frac{[NH_3]^2}{[N_2][H_2]^3}$$

An equilibrium law can be written for **homogeneous reactions** (where all the reactants and products are in the same phase) and **heterogeneous reactions** (where some reactants and products are in different phases). Because the concentrations of pure solids or liquids do not depend on how much pure substance is present, their concentrations are assigned a value of 1.

For example, at equilibrium for the heterogeneous reaction:

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

$$Q = K_c = [CO_2]$$

Both $CaCO_3$ and CaO are solids, so their concentrations are assigned a value of 1 and they do not appear in the equilibrium law expression.

## THE EXTENT OF REACTION

The value of $K_c$ gives an indication of the extent of the reaction, or the 'position of equilibrium'. Different reactions proceed to different extents.

- If $K_c < 10^{-4}$, there is a negligible forward reaction and mainly reactants present at equilibrium. The position of equilibrium is 'to the left'. (In this instance, the concentration of the initial reactants can be assumed to be the same as their final concentrations.)
- If $10^{-4} < K_c < 10^4$, the equilibrium mixture consists of significant amounts of both reactants and products.
- If $K_c > 10^4$, there is extensive forward reaction and mainly products present at equilibrium. The position of equilibrium is 'to the right'.

## FACTORS AFFECTING THE VALUE OF $K_C$

The value of the equilibrium constant for a reaction depends on:

- the coefficients in the equation: doubling the coefficients will square the original value of $K_c$, while halving the coefficients will cause $K_c$ to be the square root of the original
- the direction of the equation: reversing the equation causes $K_c$ to have the inverse value
- the temperature: the effect of a change in temperature on the value of $K_c$ is dependent on whether the reaction is exothermic or endothermic (refer back to Table 3.1.3).

As temperature increases, the equilibrium constant, $K_c$, will:

- increase for an endothermic reaction
- decrease for an exothermic reaction.

Only a change in temperature affects the value of $K_c$, and it is not changed by addition of reactants or products, pressure changes or the addition of a catalyst.

Note that the extent of reaction indicates how far the reaction has proceeded when equilibrium is achieved, while the rate of reaction indicates how fast the reaction occurred. A reaction can have a large equilibrium constant, $K_c$, but a very low rate, and vice versa.

At a specified temperature, the value of $K_c$ can be calculated from the equilibrium molar concentrations for the reactants and the products. Conversely, the concentration of a particular reactant or product can be calculated if the value of $K_c$ and the concentrations of the other reactants and products are known. The worksheets in this book will give you practice performing calculations involving equilibrium constants.

## SOLUBILITY AND EQUILIBRIUM

The expression for $K_c$ for a homogeneous equilibrium includes the molar concentrations (concentrations in mol $L^{-1}$) for all the reactants and products in the equilibrium. However, as mentioned before, for a heterogeneous equilibrium, the concentrations for a pure solid or a pure liquid are assigned the value of 1 in the expression for $K_c$. These concentrations are constant and can be removed from the expression for the $K_c$.

An example of a heterogeneous equilibrium and the associated expression for $K_c$ is the dissolution of an ionic solid in water to form an aqueous solution, e.g. for $PbI_2$:

$$PbI_2(s) \rightleftharpoons Pb^{2+}(aq) + 2I^-(aq) \qquad K_c = [Pb^{2+}][I^-]^2$$

### Measuring the solubility of ionic solids

When an ionic solid is added to water, it dissociates into separate hydrated ions. Collisions between the dissolved ions can reform the solid. Eventually, the rate of the forward reaction is equal to the rate of the reverse reaction and equilibrium is achieved. The solution becomes **saturated**.

Consider the saturated solution of solid $PbI_2$, shown in Figure 3.1.8, where equilibrium exists between the solid and its ions, $Pb^{2+}(aq)$ and $I^-(aq)$. The equation can be written as:

$$PbI_2(s) \rightleftharpoons Pb^{2+}(aq) + 2I^-(aq)$$

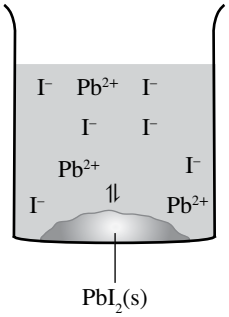


**FIGURE 3.1.8** A saturated solution of solid $PbI_2$ in equilibrium with its ions, $Pb^{2+}(aq)$ and $I^-(aq)$

Because this is a heterogeneous equilibrium system, the concentration of the solid is 1, and the equilibrium constant is written as:

$$K_{sp} = [Pb^{2+}][I^-]^2$$

An equilibrium constant for a saturated solution is called a **solubility product**, $K_{sp}$. The smaller the $K_{sp}$, the less soluble the solid is in water.

The solubility product, $K_{sp}$, of an ionic solid can be used to calculate the solubility, $s$, of the solid and equilibrium concentrations of the dissolved ions.

Solubility, $s$, is the moles per litre of the solid that has dissolved at that temperature in a saturated solution. Conversely, if the solubility is known, then the value of $K_{sp}$ can be determined. Table 3.1.4 shows the mathematical relationship between $K_{sp}$ and $s$ for two ionic solids.

You can determine whether a solution has reached equilibrium or whether more precipitation or dissolution occurs by calculating the reaction quotient, $Q$, for the system. For the lead iodide solution, $Q$ is equal to $[Pb^{2+}][I^-]^2$.



**TABLE 3.1.4** Conversion from solubility, $s$, in mol $L^{-1}$ to $K_{sp}$, where is the amount of an ionic solid that dissolves per litre

| Reaction equation | Concentration (mol $L^{-1}$) of ions in terms of solubility | $K_{sp}$ expression |
|---|---|---|
| $AgCl(s) \rightleftharpoons Ag^+(aq) + Cl^-(aq)$ | $[Ag^+] = s$ and $[Cl^-] = s$ | $K_{sp} = [Ag^+][Cl^-] = s^2$ |
| $PbI_2(s) \rightleftharpoons Pb^{2+}(aq) + 2I^-(aq)$ | $[Pb^{2+}] = s$ and $[I^-] = 2s$ | $K_{sp} = [Pb^{2+}][I^-]^2 = s \times (2s)^2 = 4s^3$ |

**TABLE 3.1.5** The relationship between $Q$, $K_{sp}$ and the position of equilibrium

| | |
|---|---|
| $Q > K_{sp}$ | Ion concentration is higher than when the system is at equilibrium. The solution is **supersaturated**. More solid forms (precipitation occurs) to establish equilibrium. |
| $Q = K_{sp}$ | Ion concentration is equal to the equilibrium value, so the system is at equilibrium. The solution is saturated. |
| $Q < K_{sp}$ | Ion concentration is lower than when the system is at equilibrium. The solution is **unsaturated**. If a solid is present, some solid dissolves to establish equilibrium. |

# Properties of acids and bases

Acids and bases are substances that have characteristic properties (Table 3.1.6).

**TABLE 3.1.6** Properties of acids and bases

| Properties of acids | Properties of bases |
|---|---|
| taste sour | taste bitter |
| turn litmus paper red | turn litmus paper blue |
| solutions conduct an electric current | solutions conduct an electric current |
| solutions have a relatively low pH | solutions have a relatively high pH |
| react with bases to produce salt and water | react with acids to produce salt and water |
| can be corrosive | can be corrosive |

Acids may be defined as proton donors, and **bases** as proton acceptors. An **acid–base reaction** involves an exchange of protons from an acid to a base.

## MONOPROTIC AND POLYPROTIC ACIDS

**Monoprotic acids** are able to donate one proton. Some acids can donate more than one proton per molecule and are referred to as **polyprotic acids** (Table 3.1.7). However, having more than one hydrogen atom in the formula does not make an acid polyprotic. For example, although ethanoic acid, $CH_3COOH$, has four hydrogens, it is a monoprotic acid because only one of its hydrogen atoms can be donated to a base. This hydrogen atom is called the **acidic proton**.

## STRENGTH OF ACIDS AND BASES

A molecule of an acid in water can donate a proton to a water molecule, which acts as a base to form a hydronium ion, $H_3O^+$ (Figure 3.1.9).

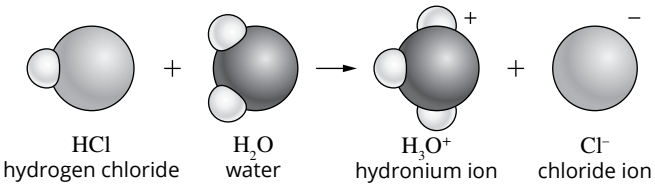


**FIGURE 3.1.9** Reaction between HCl and water. Hydrogen chloride acts as an acid by donating a proton; water acts as a base by accepting the proton.

The HCl molecule undergoes **ionisation** (because it forms ions) and **dissociation** (because it breaks apart). Acids such as HCl that dissociate completely, with all molecules donating a proton, are called **strong acids**. Acids that do not readily donate protons and incompletely dissociate in water are called **weak acids**.

Examples of the reactions of strong and weak acids are shown below.

- strong acid: $HCl(aq) + H_2O(l) \rightarrow H_3O^+(aq) + Cl^-(aq)$
- weak acid: $CH_3COOH(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + CH_3COO^-(aq)$

Soluble ionic compounds that contain the $OH^-$ ion, such as NaOH, dissociate in water and release the $OH^-$ ion, producing a strongly basic solution.

Some molecules, such as ammonia, act as a base in water. To a limited extent, ammonia accepts a proton from a water molecule and produces a hydroxide ion, $OH^-$ (Figure 3.1.10). The molecules can be said to ionise (because they form ions) and dissociate (because they break up).

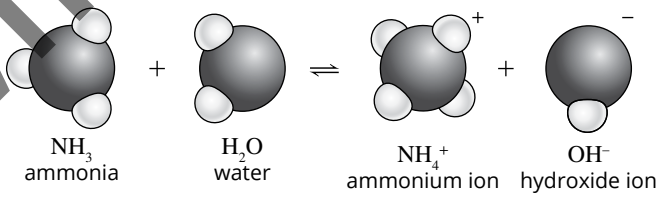


**FIGURE 3.1.10** Reaction between ammonia and water: the ammonia acts as a base and accepts a proton

Like acids, bases can be classed as strong or weak. Table 3.1.8 summarises the differences between the strengths of acids and bases.

**TABLE 3.1.7** Monoprotic and polyprotic acids

| Type of acid | Number of protons to donate | Example | Equation |
|---|---|---|---|
| monoprotic acid | 1 | nitric acid, $HNO_3$, ethanoic acid, $CH_3COOH$ | $HNO_3(aq) + H_2O(l) \rightarrow NO_3^-(aq) + H_3O^+(aq)$<br>$CH_3COOH(aq) + H_2O(l) \rightleftharpoons CH_3COO^-(aq) + H_3O^+(aq)$ |
| diprotic acid | 2 | sulfuric acid, $H_2SO_4$ | $H_2SO_4(aq) + H_2O(l) \rightarrow HSO_4^-(aq) + H_3O^+(aq)$<br>$HSO_4^-(aq) + H_2O(l) \rightleftharpoons SO_4^{2-}(aq) + H_3O^+(aq)$ |
| triprotic acid | 3 | phosphoric acid, $H_3PO_4$ | $H_3PO_4(aq) + H_2O(l) \rightleftharpoons H_2PO_4^-(aq) + H_3O^+(aq)$<br>$H_2PO_4^-(aq) + H_2O(l) \rightleftharpoons HPO_4^{2-}(aq) + H_3O^+(aq)$<br>$HPO_4^{2-}(aq) + H_2O(l) \rightleftharpoons PO_4^{3-}(aq) + H_3O^+(aq)$ |



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**TABLE 3.1.8** Strong and weak acids and bases

| | Acids | Bases |
|---|---|---|
| **Strong** | readily donate protons and completely dissociate in water (Figure 3.1.11a)<br>$HCl(aq) + H_2O(l) \rightarrow H_3O^+(aq) + Cl^-(aq)$ | readily accept protons and completely dissociate in water<br>$Na_2O(s) + H_2O(l) \rightarrow 2Na^+(aq) + OH^-(aq)$<br>$NaOH(s) \xrightarrow{H_2O(l)} Na^+(aq) + OH^-(aq)$ |
| **Weak** | do not readily donate protons and incompletely dissociate in water (Figure 3.1.11b)<br>$CH_3COOH(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + CH_3COO^-(aq)$ | do not readily accept protons and incompletely dissociate in water (Figure 3.1.11c)<br>$NH_3(aq) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$ |

Figure 3.1.11 illustrates the differences between strong and weak acids and bases.

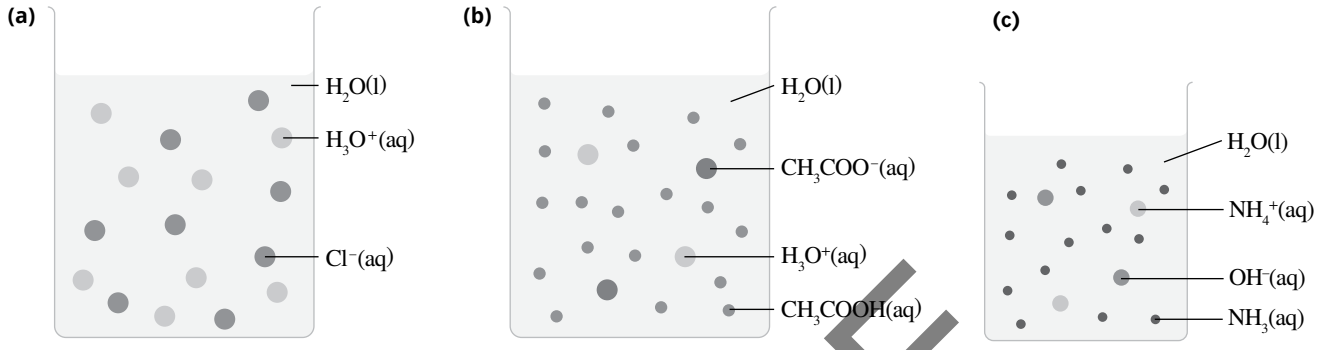



**FIGURE 3.1.11** (a) In a 0.1 mol $L^{-1}$ solution of HCl, the acid molecules dissociate completely in water. HCl is a strong acid. (b) In a 0.1 mol $L^{-1}$ solution of $CH_3COOH$, only a small proportion of the $CH_3COOH$ molecules dissociate in water. $CH_3COOH$ is a weak acid. (c) Only a small proportion of $NH_3$ molecules in an ammonia solution are dissociated at any given time. $NH_3$ is a weak base.

Measurements of electrical conductivity can be used to distinguish between strong and weak acids and to determine the extent of their dissociation. The $OH^-$ and $H_3O^+$ ions are small and mobile, and readily conduct electricity. The concentration of these ions therefore determines the conductivity of solutions of equal acid concentrations. Figure 3.1.12 shows that 0.1 mol $L^{-1}$ HCl is a much better conductor than 0.1 mol $L^{-1}$ $CH_3COOH$.

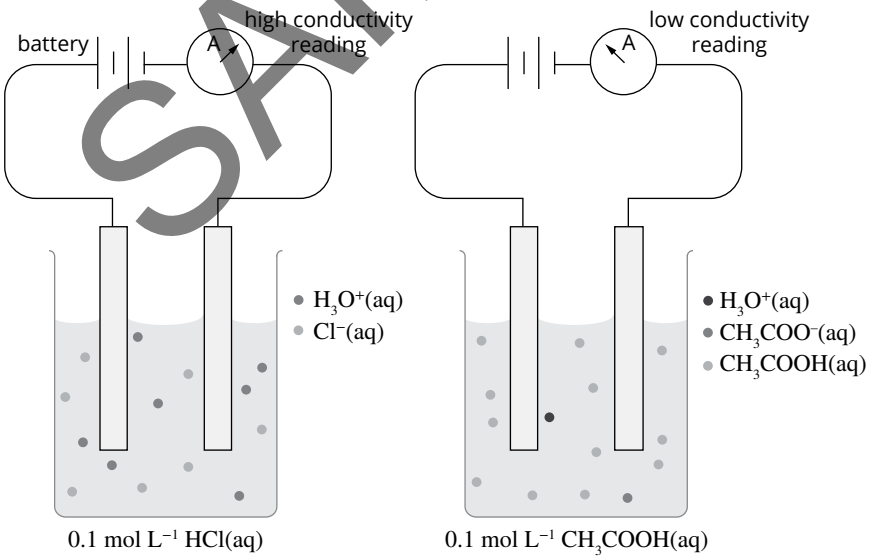



**FIGURE 3.1.12** Electrical conductivity of 0.1 mol $L^{-1}$ solutions of a strong acid, HCl, and a weak acid, $CH_3COOH$

$CH_3COOH$ is a weak acid and therefore only partially dissociates in an aqueous solution, producing fewer mobile ions. HCl is a strong acid that almost completely dissociates in aqueous solution.

## Strength versus concentration

Acids and bases can be described as either strong or weak, and also as **concentrated** or **dilute** (Figure 3.1.13 on page 12). The strength of an acid or a base is determined by its ability to dissociate in solution, whereas the concentration of an acid or base refers to the amount of acid or base present in a given volume of solution.

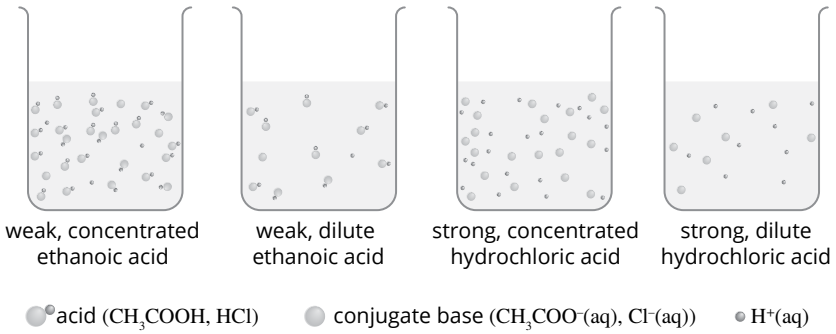



**FIGURE 3.1.13** The concentration of ions in an acid solution depends on both the concentration and the strength of the acid.

The concentration of an acid or base (or any solute in solution) is calculated using the formula:

$$c = \frac{n}{V}$$

where $c$ is the concentration in mol $L^{-1}$, $n$ is the amount in mol of the solute and $V$ is the volume of the solution in L.

When solutions are diluted, the number of moles of solute does not change, but the volume of water is increased, which causes the concentration to decrease. A useful formula for calculating the concentration of a diluted solution is:

$$c_1 \times V_1 = c_2 \times V_2$$

where $c_1$ and $V_1$ are the initial concentrations and volume, and $c_2$ and $V_2$ are the final concentrations and volume.

# pH scale

The **acidity** of a solution is conveniently measured using a quantity called **pH**. The pH of a solution is a function of the concentration of hydronium ions present in the solution:

$$pH = -\log_{10}[H_3O^+]$$

This equation can be rearranged and written as:

$$[H_3O^+] = 10^{-pH}$$

The pH scale is shown in Figure 3.1.14. As you can see, the more acidic a solution becomes, the lower its pH. The more basic a solution becomes, the higher its pH. In general, at 25°C:

- pH < 7 is an acidic solution
- pH = 7 is a neutral solution
- pH > 7 is a basic solution.

The **basicity** of a solution can be measured using the quantity called **pOH**. The pOH of a solution is a function of the concentration of hydroxide ions present in the solution:

$$pOH = -\log_{10}[OH^-]$$

This equation can be rearranged and written as:

$$[OH^-] = 10^{-pOH}$$

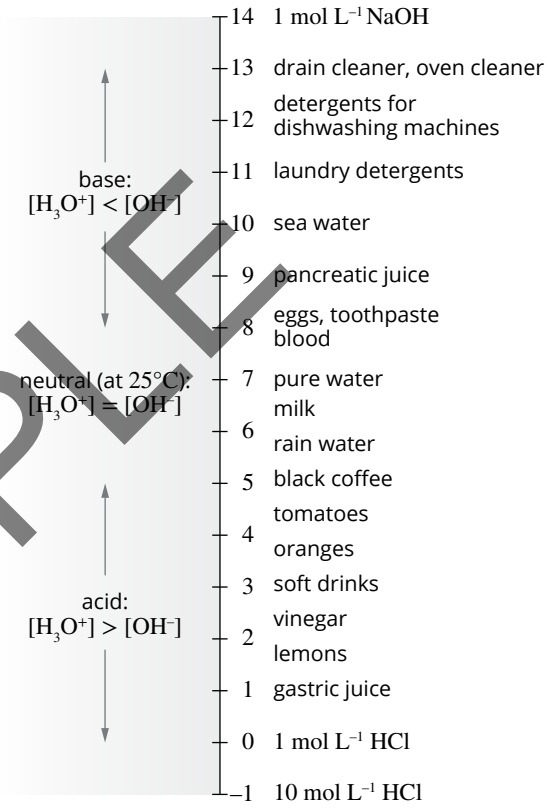



**FIGURE 3.1.14** The pH scale and the pH of solutions of some common household substances

## IONIC PRODUCT OF WATER ($K_W$)

In an aqueous solution, water molecules react to a very slight extent to form $H_3O^+$ and $OH^-$ ions:

$$H_2O(l) + H_2O(l) \rightleftharpoons H_3O^+(aq) + OH^-(aq)$$
$$\Delta H = +57.6\,kJ\,mol^{-1}$$

This reaction is called the **self-ionisation** of water, and an equilibrium constant can be written as:

$$K_c = \frac{[H_3O^+][OH^-]}{[H_2O]^2}$$

Because the concentration of water in aqueous solutions, $[H_2O]$, is approximately $55\,mol\,L^{-1}$ and relatively constant, it is convenient to include the $[H_2O]^2$ term as part of the equilibrium constant. So,

$$K_c \times [H_2O]^2 = K_w = [H_3O^+][OH^-]$$
$$= 1.0 \times 10^{-14} \text{ at } 25°C$$

where $K_w$ is called the **ionic product of water** (also called the ionisation constant of water).

A simple relationship between pH and pOH can also be derived by taking the negative logarithm of both sides of the expression for $K_w$. At 25°C, we obtain the relationship:

$$pH + pOH = 14$$

Since $K_w = [H_3O^+][OH^-]$, you can see that as $[H_3O^+]$ increases, the $[OH^-]$ decreases, and vice versa. If the concentration of $OH^-$ in a solution is known, the concentration of $H_3O^+$ in the solution at 25°C can be calculated from the expression $[H_3O^+][OH^-] = 1.0 \times 10^{-14}$, and vice versa.

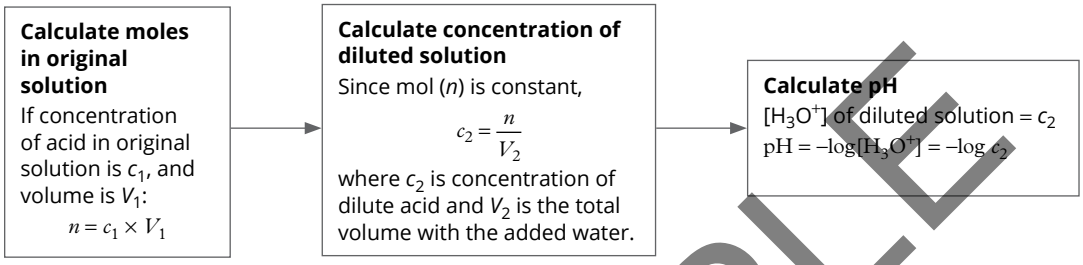
In pure water at 25°C, $[H_3O^+] = [OH^-]$, so the concentration of both $H_3O^+$ and $OH^-$ is $1.0 \times 10^{-7}$ mol $L^{-1}$.

Because $K_w$ is an equilibrium constant, its value is temperature-dependent. The ionic product of water is $1.0 \times 10^{-14}$ only at 25°C. Because the self-ionisation reaction between water molecules is an endothermic reaction, we know from Le Châtelier's principle that as the temperature increases, $K_w$ will also increase. As a consequence, $[H_3O^+]$ will increase and the pH of pure water will decrease as the temperature increases.

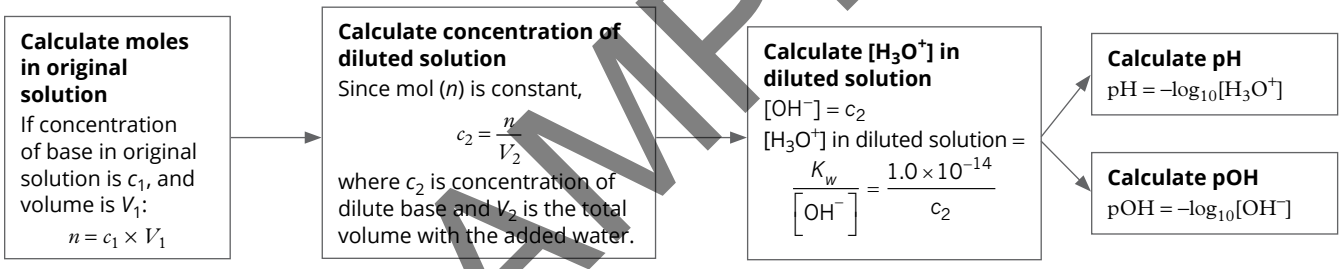
## CALCULATING pH FOR DILUTIONS AND MIXTURES

The following flow charts (Figure 3.1.15) show how to calculate pH when a strong acid or a strong base is diluted, or when a strong acid and a strong base are mixed together. The worksheets in this book will give you practice with these calculations.

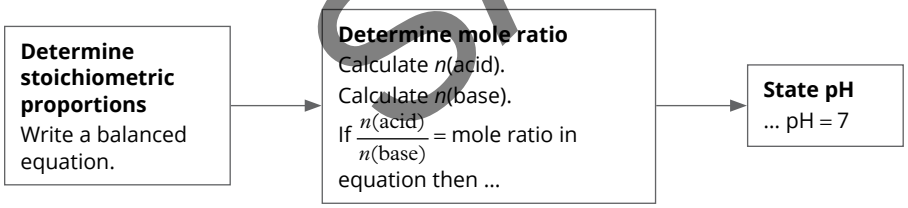
**(a)** Dilution of a strong monoprotic acid




**(b)** Dilution of a strong base




**(c)** Reaction of a strong acid and a strong base in stoichiometric proportions




**(d)** Reaction of a strong acid and a strong base where the acid is in excess

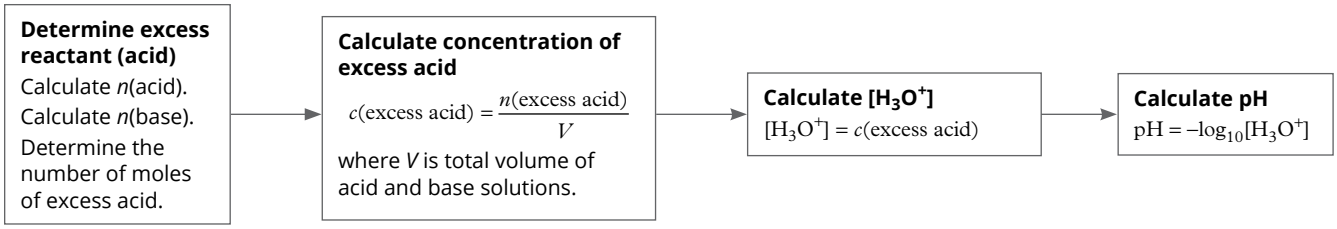



**FIGURE 3.1.15** Calculating pH for (a) dilution of a strong monoprotic acid, (b) dilution of a strong base, (c) the reaction of a strong acid and a strong base in stoichiometric proportions and (d) the reaction of a strong acid and a strong base where the acid is in excess

# Brønsted–Lowry model

## CONJUGATE PAIRS

The Brønsted–Lowry theory describes an acid–base reaction as a reaction in which proton transfer occurs. This reaction can be represented by either a formula or an ionic equation.

In an acid–base reaction, the acid donates a proton to form its **conjugate base,** and the base accepts a proton and becomes its **conjugate acid.** A **conjugate acid–base pair** is two species that differ by a single $H^+$ ion. For example:

$$\underset{\text{acid}}{HCl(aq)} + \underset{\text{base}}{OH^-(aq)} \rightarrow \underset{\text{conjugate base}}{Cl^-(aq)} + \underset{\text{conjugate acid}}{H_2O(l)}$$

$Cl^-$ is the conjugate base of HCl, and $H_2O$ is the conjugate acid of $OH^-$. Table 3.1.9 shows various conjugate pairs in some acid–base reactions.

## AMPHIPROTIC SUBSTANCES

An **amphiprotic** substance is one that can act as either an acid or a base, depending on the substance with which it is reacting. In other words, it is able to donate or accept protons. The $HCO_3^-$, $HSO_4^-$ and $H_2PO_4^-$ ions are examples of amphiprotic substances, as shown in Table 3.1.10.

**TABLE 3.1.9** Conjugate pairs in some acid–base reactions

| Ionic equation | Conjugate pairs |
|---|---|
| $NH_3(aq) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$ | $NH_3$ and $NH_4^+$<br>$H_2O$ and $OH^-$ |
| $HCl(aq) + OH^-(aq) \rightleftharpoons Cl^-(aq) + H_2O(l)$ | HCl and $Cl^-$<br>$OH^-$ and $H_2O$ |
| $H_2SO_4(aq) + H_2O(l) \rightleftharpoons HSO_4^-(aq) + H_3O^+(aq)$ | $H_2SO_4$ and $HSO_4^-$<br>$H_2O$ and $H_3O^+$ |
| $CH_3COOH(aq) + OH^-(aq) \rightleftharpoons CH_3COO^-(aq) + H_2O(l)$ | $CH_3COOH$ and $CH_3COO^-$<br>$OH^-$ and $H_2O$ |

**TABLE 3.1.10** Amphiprotic substances

| Ionic equation | Conjugate pairs |
|---|---|
| $HCO_3^-$ acting as an acid: $HCO_3^-(aq) + H_2O(l) \rightleftharpoons CO_3^{2-}(aq) + H_3O^+(aq)$<br>$HCO_3^-$ acting as a base: $HCO_3^-(aq) + H_2O(l) \rightleftharpoons H_2CO_3(aq) + OH^-(aq)$ | $HCO_3^-$ and $CO_3^{2-}$, $H_2O$ and $H_3O^+$<br>$HCO_3^-$ and $H_2CO_3$, $H_2O$ and $OH^-$ |
| $HSO_4^-$ acting as an acid: $HSO_4^-(aq) + H_2O(l) \rightleftharpoons SO_4^{2-}(aq) + H_3O^+(aq)$<br>$HSO_4^-$ acting as a base: $HSO_4^-(aq) + H_2O(l) \rightleftharpoons H_2SO_4(aq) + OH^-(aq)$ (Note that, in practice, this does not occur to any extent because $H_2SO_4$ is a strong acid.) | $HSO_4^-$ and $SO_4^{2-}$, $H_2O$ and $H_3O^+$<br>$HSO_4^-$ and $H_2SO_4$, $H_2O$ and $OH^-$ |
| $H_2PO_4^-$ acting as an acid: $H_2PO_4^-(aq) + H_2O(l) \rightleftharpoons HPO_4^{2-}(aq) + H_3O^+(aq)$<br>$H_2PO_4^-$ acting as a base: $H_2PO_4^-(aq) + H_2O(l) \rightleftharpoons H_3PO_4(aq) + OH^-(aq)$ | $H_2PO_4^-$ and $HPO_4^{2-}$, $H_2O$ and $H_3O^+$<br>$H_2PO_4^-$ and $H_3PO_4$, $H_2O$ and $OH^-$ |



## BUFFERS

**Buffer solutions** are solutions that contain appreciable amounts of a weak acid in equilibrium with its conjugate weak base. Buffers resist changes to pH.

The most effective buffers are prepared by mixing approximately equal amounts of a weak acid with a salt of its conjugate base. To make a buffer of a particular pH, the conjugate acid–base pair that is chosen should have a $pK_a$ of the acid as close as possible to the desired pH.

Table 3.1.11 describes how a buffer solution made from ethanoic acid and sodium ethanoate resists a change in pH change when HCl is added. The buffer contains the acid and its conjugate base in equilibrium:

$$CH_3COOH(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + CH_3COO^-(aq)$$

**TABLE 3.1.11** How a buffer consisting of ethanoic acid and ethanoate ions resists change to pH

| Stage | Explanation |
|---|---|
| initial | Equilibrium between buffer acid and buffer base. Equal amounts of both the acid and base are present in the buffer. |
| addition of a small amount of HCl | As Le Châtelier's principle predicts, the added $H^+(aq)$ causes a net reverse reaction; the reaction between the added acid and buffer base produces more buffer acid. |
| final | Equilibrium is restored between the buffer acid and buffer base. Most of the added $H^+(aq)$ has been neutralised, and there is a little change in pH. |

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Conversely, the addition of a small amount of a strong base will favour the forward reaction. Most of the added $OH^-(aq)$ will be neutralised by reaction with the buffer acid, and again, there will be a little change in pH.

# Dissociation constants

## ACIDITY CONSTANT, $K_a$

Consider the dissociation of the weak acid, ethanoic acid, in water:

$$CH_3COOH(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + CH_3COO^-(aq)$$

An equilibrium constant can be written as:

$$K_c = \frac{[H_3O^+][CH_3COO^-]}{[CH_3COOH][H_2O]}$$

Because $[H_2O]$ is almost constant, you can write:

$$K_c[H_2O] = K_a = \frac{[H_3O^+][CH_3COO^-]}{[CH_3COOH]}$$

$$= 1.75 \times 10^{-5} \text{ at } 25^oC$$

The constant $K_a$ is called the **acidity constant** (or acid dissociation constant) of ethanoic acid. An acidity constant is a measure of the strength of an acid.

The small value of $K_a$ for $CH_3COOH$ indicates that the extent of the reaction is limited and that $CH_3COOH$ is a weak acid. By comparison, the $K_a$ for HCl is $10^7$ at 25°C. The reaction of HCl with water is almost complete at equilibrium, and HCl is classed as a strong acid.

## BASICITY CONSTANT, $K_b$

Similarly, the **basicity constant** (or base dissociation constant), $K_b$, of a base is used to measure the strength of the base. For example, the conjugate base of ethanoic acid is the ethanoate ion, $CH_3COO^-(aq)$, and it reacts in water as shown:

$$CH_3COO^-(aq) + H_2O(l) \rightleftharpoons CH_3COOH(aq) + OH^-(aq)$$

$$K_b = \frac{[CH_3COOH][OH^-]}{[CH_3COO^-]}$$

$$= 5.7 \times 10^{-10} \text{ at } 25^oC$$

The following relationship can be written for any conjugate pair:

$$K_a \times K_b = K_w$$

Common acids and bases and their $K_a$ and $K_b$ values are listed in Table 3.1.12.



**TABLE 3.1.12** Common strong and weak acids and bases, with dissociation constants measured at 25°C

| | Acids | $K_a$ | p$K_a$ | Bases | $K_b$ | p$K_b$ |
|---|---|---|---|---|---|---|
| **Strong** | HCl<br>$HNO_3$<br>$H_2SO_4$ | $1 \times 10^7$<br>$2 \times 10^1$<br>$1 \times 10^{9*}$ | –7<br>–1.3<br>–9 | All group 1 hydroxides and $Ba(OH)_2$ | >1 | <0 |
| **Weak** | All carboxylic acids, e.g. $CH_3COOH$<br>$H_2CO_3$ | approx. $10^{-5}$<br>$4.5 \times 10^{-7*}$ | approx. 5<br>6.3 | $NH_3$<br>Organic amines, e.g. $CH_3NH_2$ | $1.8 \times 10^{-5}$<br>approx. $10^{-3}$ | 4.7<br>approx. 3 |

*$K_a$ for the loss of the first proton of $H_2SO_4$ and $H_2CO_3$

Just as we defined pH as $-\log_{10}[H_3O^+]$, so we can write:

$$pK_a = -\log_{10}K_a$$
$$pK_b = -\log_{10}K_b$$

Table 3.1.13 summarises the key formulas you need to know for your study of acids and bases.

**TABLE 3.1.13** Acid–base formulas

| | | | |
|---|---|---|---|
| $pH = -\log_{10}[H_3O^+]$ | $[H_3O^+] = 10^{-pH}$ | pH + pOH = 14 (at 25°C) | $[H_3O^+] \times [OH^-] = 1.00 \times 10^{-14}$ (at 25°C) |
| $pOH = -\log_{10}[OH^-]$ | $[OH^-] = 10^{-pOH}$ | | |
| $pK_a = -\log_{10}K_a$ | $K_a = 10^{-pK_a}$ | $pK_a + pK_b = 14$ (at 25°C) | $K_a \times K_b = K_w = 1.00 \times 10^{-14}$ (at 25°C) |
| $pK_b = -\log_{10}K_b$ | $K_b = 10^{-pK_b}$ | | |

The dissociation reaction with water for a general acid HA and its conjugate base $A^-$ is:

$$HA(aq) + H_2O(l) \rightarrow A^-(aq) + H_3O^+(aq)$$

$$K_a = \frac{[H_3O^+][A^-]}{[HA]}$$

The dissociation reaction with water for a general base B and its conjugate acid $BH^+$ is:

$$B(aq) + H_2O(l) \rightarrow BH^+(aq) + OH^-(aq)$$

$$K_b = \frac{[BH^+][OH^-]}{[B]}$$

The worksheets in this book will give you practice in calculations involving acids and bases.

# Acid–base indicators

An acid–base **indicator** is a weak acid that has a conjugate base of a different colour. The colour of an indicator in solution depends on the relative concentrations of the acid and base, and therefore depends on the pH of the solution. Table 3.1.14 lists some common indicators and their pH range (the pH over which they change colour).

**TABLE 3.1.14** Some indicators with their p$K_a$ values and pH ranges

| Indicator | p$K_a$ | Colour in acid | Colour in base | pH range |
|---|---|---|---|---|
| phenolphthalein | 9.3 | colourless | pink | 8.2–10.0 |
| methyl orange | 3.4 | pink | yellow | 3.2–4.4 |
| bromothymol blue | 7.1 | yellow | blue | 6.0–7.6 |

Indicators undergo reversible reactions between their acid and base forms. If the acid form is represented by the general formula HIn and the base form as $In^-$, the reaction can be represented as:

$$HIn(aq) + H_2O(l) \rightleftharpoons In^-(aq) + H_3O^+(aq)$$

with an acidity constant given by:

$$K_a = \frac{[H_3O^+][In^-]}{HIn}$$

When $[In^-] = [HIn]$, the $K_a = [H_3O^+]$, so $pK_a = pH$. This occurs when the indicator changes colour. The colour change can be assumed to occur within a pH range equal to $pK_a \pm 1$. Table 3.1.14 includes the $pK_a$ values for the indicators shown.

Indicators are used in acid–base titrations to indicate the **equivalence point** of the reaction. The equivalence point is the point at which stoichiometric amounts of reactants have been mixed. If an indicator is selected carefully, the **end point**—the point at which the indicator changes colour—and the equivalence point of the reaction will occur at nearly the same pH.

# Volumetric analysis

Acid–base titrations are a type of volumetric analysis that can be used to determine the concentration of an acid or a base in water. In a simple acid–base titration (Figure 3.1.16), one solution is dispensed from a **burette** into a conical flask that contains a volume of another solution, which has been precisely measured using a **pipette**. These solutions are usually colourless, so an acid–base indicator is added to the conical flask. The indicator selected must be one that changes colour (at its 'end point') when the equivalence point has been reached, making it possible to see when the reaction is complete.



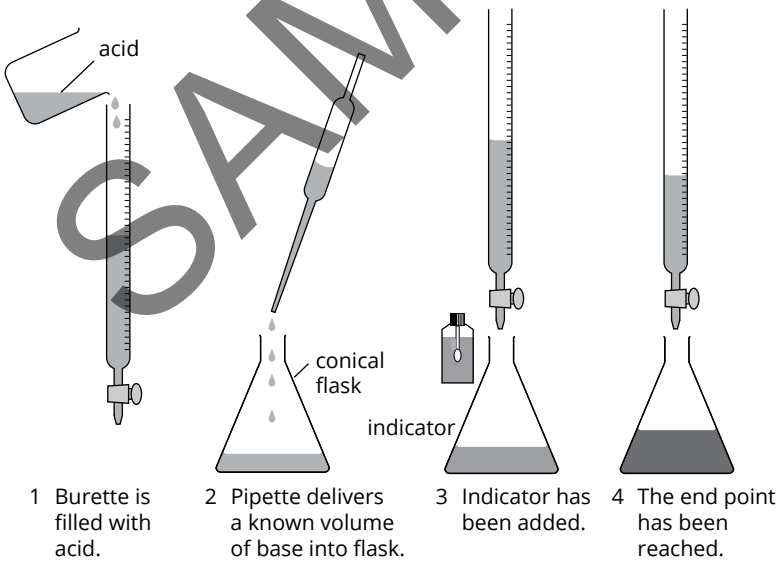


**FIGURE 3.1.16** Steps in a titration

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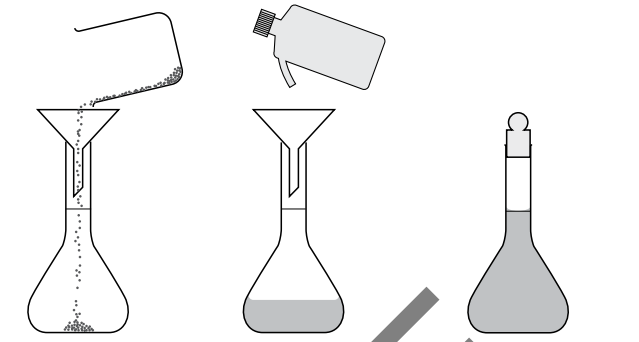
## TERMINOLOGY IN VOLUMETRIC ANALYSIS

Some of the terms used in volumetric analysis are explained in Table 3.1.15.

**TABLE 3.1.15** Terms used in volumetric analysis

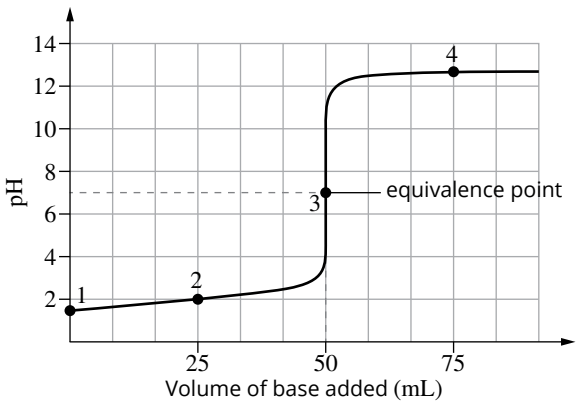
| Term | Description |
|---|---|
| primary standard | A substance so pure that the amount can be accurately calculated from its mass. For a chemical to be a primary standard, it should be easily obtained, have a known formula and be easily stored without reacting with water or gases in the atmosphere. $Na_2CO_3$ is an example of a primary standard. |
| standard solution | A solution with an accurately known concentration. This diagram shows a standard solution being made from a primary standard.   |
| burette | a calibrated glass tube that delivers variable volumes accurately |
| titre | the volume delivered by a burette |
| pipette | a calibrated glass tube that delivers a fixed volume accurately |
| aliquot | volume delivered by a pipette |
| end point | the point at which the indicator changes colour |
| equivalence point | the point in the titration where the stoichiometric proportions of the reactants have been mixed |
| indicator | a weak acid that has a conjugate base of a different colour |
| analyte | a substance whose chemical constituents are being identified and measured |

Instead of using an indicator, a titration can be monitored using a pH meter.

## TITRATION CURVES

If a pH meter is used for a titration, a **titration curve** (also known as a pH curve) can be plotted. This is a graph of pH against volume of base (or acid) added. The equivalence point can be identified from the shape of the curve. Table 3.1.16 shows the titration curves for different reactions with explanations of the pH at different points along the curves. The equivalence point occurs when the gradient of the pH curve is steepest.

**TABLE 3.1.16** Explanation of the shape of pH curves

| Reaction | Titration curve | Stages | pH | Explanation of shape and pH |
|---|---|---|---|---|
| strong acid, HCl, with strong base, NaOH |   | 1 | very low | only acid present initially; high $[H_3O^+]$ due to complete ionisation of strong acid |
| | | 2 | low | half of the $H_3O^+$ ions have been neutralised by the addition of $OH^-$ ions |
| | | 3 | 7 | equivalence point, where initial moles of HCl equals moles of NaOH added; $[H_3O^+] = [OH^-]$; neutral salt and water are products; $[H_3O^+] = [OH^-] = 1 \times 10^{-7}$ mol $L^{-1}$ |
| | | 4 | very high | excess $OH^-$ ions as more base is added |

*continued over to next page*

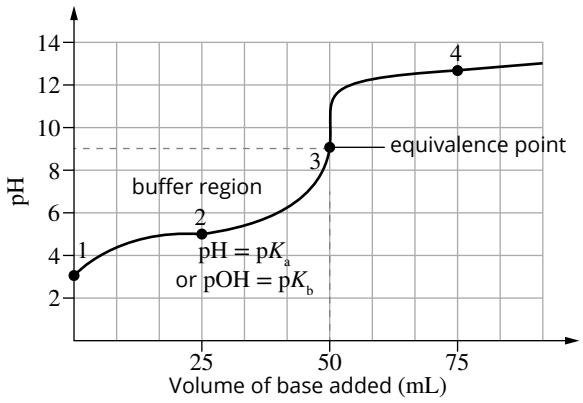
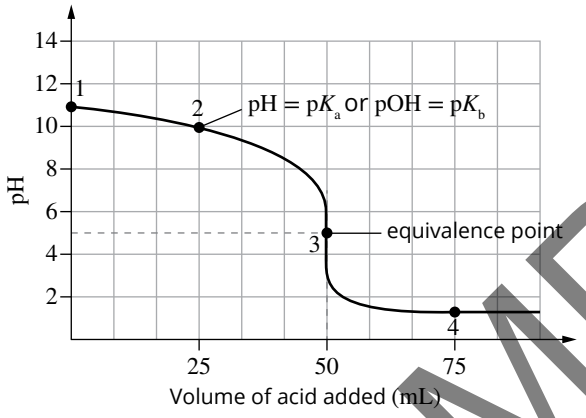
| Reaction | Titration curve | Stages | pH | Explanation of shape and pH |
|---|---|---|---|---|
| weak acid, $CH_3COOH$, with strong base, NaOH |  | 1 | low | only acid present initially; pH below 7, due to partial ionisation of weak acid |
| | | 2 | equal to p$K_a$ of acid, $CH_3COOH$ pOH = p$K_b$ of conjugate base, $CH_3COO^-$ | half of the weak acid has been neutralised by the addition of $OH^-$, producing the weak base $CH_3COO^-$; since $[CH_3COOH] = [CH_3COO^-]$, pH = p$K_a$; this is called the buffer region |
| | | 3 | >7 | equivalence point, where initial moles of $CH_3COOH$ equals moles of NaOH added; the conjugate base of the weak acid $CH_3COO^-$ is formed, which ionises slightly to produce $OH^-$, so the pH is more than 7 |
| | | 4 | very high | excess $OH^-$ ions as more base is added |
| weak base, $NH_3$, with strong acid, HCl |  | 1 | high | only base present initially; pH above 7, due to partial ionisation of weak base |
| | | 2 | equal to p$K_a$ of conjugate acid, $NH_4^+$ pOH = p$K_b$ of base, $NH_3$ | half of the $H_3O^+$ ions have been neutralised by the addition of $NH_3$, producing $NH_4^+$, so $[NH_3] = [NH_4^+]$; this is called the buffer region |
| | | 3 | <7 | equivalence point, where moles of HCl equals initial moles of $NH_3$; the conjugate acid of the weak base $NH_4^+$ is formed, which ionises slightly to produce $H_3O^+$, so the pH is less than 7 |
| | | 4 | very low | excess $H_3O^+$ ions as more acid is added |



Figure 3.1.17 compares the titration curves of **(a)** a strong acid and a strong base with a strong acid and a weak base, and **(b)** a strong acid and a strong base with a weak acid and a strong base. Note that:

- for a strong acid and a strong base titration, pH = 7 at the equivalence point
- for a strong acid and a weak base titration, pH < 7 at the equivalence point
- for a weak acid and a strong base titration, pH > 7 at the equivalence point.

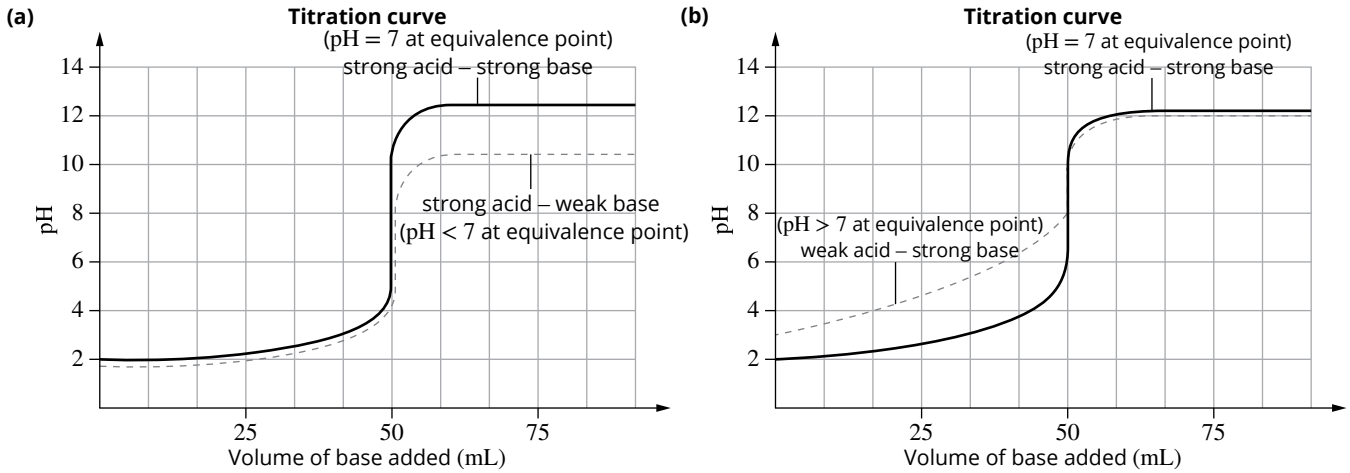



**FIGURE 3.1.17** (a) Titration curves for a strong acid and a strong base with a strong acid and a weak base, and (b) a strong acid and a strong base with a weak acid and a strong base

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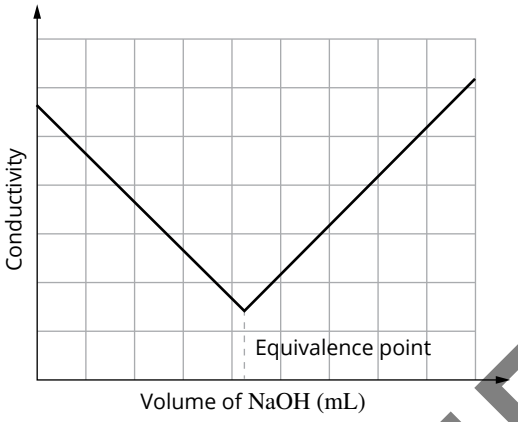
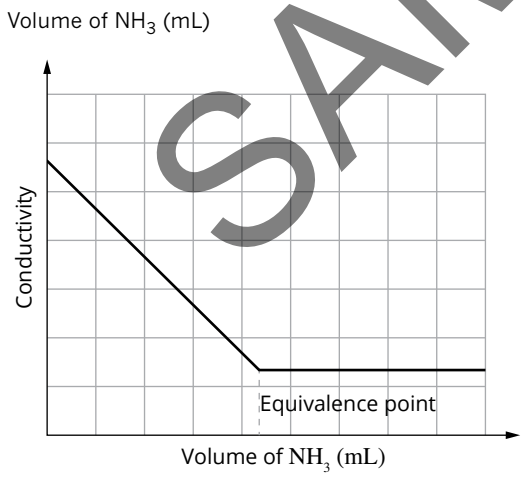
# Conductometric titrations

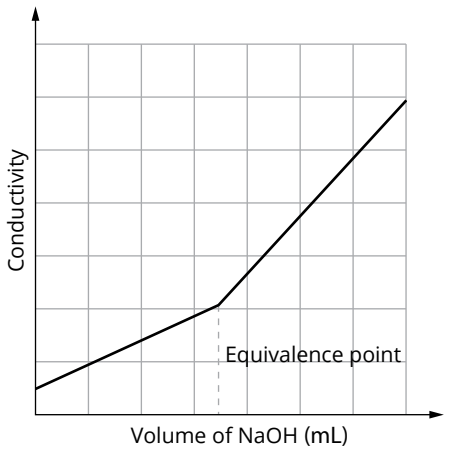
An alternative to using an indicator in a titration to determine an equivalence point is to measure the conductivity of the solution in the flask. When titrating a strong acid with a strong base, a graph of the conductivity of the acid solution against the volume of base added shows a sharp dip at the equivalence point.

The conductivity of the solution is largely dependent on the concentration of $H_3O^+$ and $OH^-$ ions, which are small and mobile. Table 3.1.17 shows the conductometric titration curves that are observed for titrations of different combinations of strong and weak acids and bases, and explains the reasons for the shapes of the graphs. The equivalence point occurs at the change in gradient of the curve, where the two straight lines intersect, and is indicated by the dashed grey line in the graphs below.

**TABLE 3.1.17** Conductometric titration curves and explanations for their shapes

| Reaction | Conductivity graph | Stages | Concentration of mobile ions, $H_3O^+$ and $OH^-$ | Conductivity |
|---|---|---|---|---|
| strong acid, HCl, with strong base, NaOH |  | initially | high $[H_3O^+]$ | high |
| | | gradual addition of base | decrease in $[H_3O^+]$ due to neutralisation by $OH^-$ ions | decreases |
| | | at equivalence point | $[H_3O^+] = [OH^-] = 10^{-7}$ mol $L^{-1}$ (a low concentration) (the reaction mixture is simply a NaCl solution) | at a minimum; only larger, slow-moving ions are present |
| | | continued addition of base after equivalence point | high $[OH^-]$ | increases |
| strong acid, HCl, with weak base, $NH_3$ |  | initially | high $[H_3O^+]$ | high |
| | | gradual addition of base | decrease in $[H_3O^+]$ due to neutralisation by the added base | decreases |
| | | at equivalence point | almost all the $H_3O^+$ has reacted (the reaction mixture is simply a $NH_4Cl$ solution) | at a minimum; mainly larger, slow-moving ions are present |
| | | continued addition of base after equivalence point | low $[H_3O^+]$ and $[OH^-]$; almost no more ions being added | remains approximately constant because of the partial ionisation of the weak base and minimal concentration of mobile $OH^-$ ions |

*continued over to next page*

| Reaction | Conductivity graph | Stages | Concentration of mobile ions, $H_3O^+$ and $OH^-$ | Conductivity |
|---|---|---|---|---|
| weak acid, $CH_3COOH$, with strong base, NaOH |   | initially | low [$H_3O^+$] | low; the partial ionisation of the weak acid produces a low concentration of mobile $H_3O^+$ ions |
| | | gradual addition of base | neutralises weak acid to produce less mobile conjugate base ions, $CH_3COO^-$. | gradually increases due to increasing concentration of $Na^+$ and the conjugate base ions |
| | | at equivalence point | the reaction mixture is simply a $NaCH_3COO$ solution; low concentrations of $H_3O^+$ and $OH^-$ | significant because of a high concentration of $Na^+$ and the conjugate base ions |
| | | continued addition of base after equivalence point | high [$OH^-$] | increases |

## CALCULATIONS USING VOLUMETRIC ANALYSIS DATA

Volumetric analysis involves performing volume–volume stoichiometry calculations. To calculate an unknown concentration of an acid or a base using volume–volume stoichiometry, you need to know the volumes of both reactants in the titration, as well as the concentration of one reactant. The calculation requires a balanced equation, the use of stoichiometric ratios and the formula $n = c \times V$ (Figure 3.1.18). The worksheets in this book provide you with practice in performing these calculations.

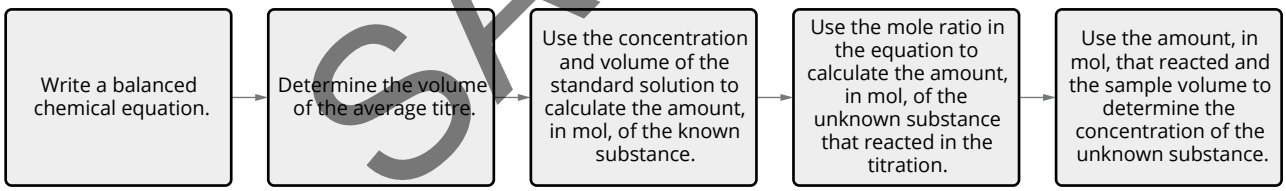



**FIGURE 3.1.18** Flow chart summarising the steps in the calculation of the concentration of an unknown substance using data from a titration



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## WORKSHEET 3.1.1

# Knowledge preview—thinking about rates and energy

**1** Identify the correct term for each definition by completing the following table. This will help you check your knowledge and understanding of the key ideas involved in rates of reaction and energy of reaction, in preparation for your study of equilibrium.

| Definition | Correct term |
|---|---|
| The energy needed to break the bonds between atoms in the reactants to enable a reaction to occur. | |
| The name for a theory that accounts for the rates of chemical reactions. | |
| A substance that increases the rate of reaction but is not consumed in the reaction. | |
| A chemical reaction in which all the species are in the same phase. | |
| A reaction that releases energy to the surroundings. | |
| The name for a chemical equation that includes the enthalpy change of the reaction. | |
| A reaction for which the $\Delta H$ is positive. | |

**2** Show the balanced equations for the reactions that occur when the following substances dissolve in water. Include states of matter in the equations.

HCl(g) _____

NaOH(s) _____

$NH_3(g)$ _____

$CH_3COOH(l)$ _____

**3** Calculate the concentration of a standard solution prepared by dissolving 3.65 g of sodium carbonate, $Na_2CO_3$, in water and making the volume up to 250.0 mL in a volumetric flask.

**4** Determine the mass of sodium carbonate, $Na_2CO_3$, you would need to dissolve in a 500 mL volumetric flask to make up a 1.00 mol $L^{-1}$ solution.

# PRACTICAL ACTIVITY 3.1.1

# Effect of concentration changes on equilibrium yields

**Suggested duration:** 50 minutes

## RESEARCH QUESTION

What is the effect of concentration changes on an aqueous equilibrium?

## RATIONALE

You will be given a solution of $Fe(SCN)^{2+}$ that contains the ions $Fe^{3+}$, $SCN^-$ and $Fe(SCN)^{2+}$ at equilibrium:

$$Fe^{3+}(aq) + SCN^-(aq) \rightleftharpoons Fe(SCN)^{2+}(aq)$$

pale yellow colourless red

The intense, blood-red colour of the solution is due to the presence of the $Fe(SCN)^{2+}$ ion. The colour of the solution in each test tube, when viewed down the tube, is a measure of the amount of $Fe(SCN)^{2+}$ ions, in mol, present in the tube.

By noting how the intensity of this colour changes, it is possible to deduce the effect of each of the tests performed in this experiment on the equilibrium.

For instance, if the colour deepens, the number of $Fe(SCN)^{2+}$ ions has increased and the number of $Fe^{3+}$ and $SCN^-$ ions must have simultaneously decreased because they are used up to form more $Fe(SCN)^{2+}$. The equilibrium would be described as having a net forward reaction (its position would have 'shifted to the right').

In Tests A–E, the amount (number of moles) of $Fe^{3+}$ or $SCN^-$ ions present in the solution is initially changed as follows.

- In Test A, addition of $Fe(NO_3)_3(aq)$ increases the amount of $Fe^{3+}$.
- In Test B, addition of KSCN(aq) increases the amount of $SCN^-$.
- In Test C, addition of NaF(aq) decreases the amount of $Fe^{3+}$ because $F^-$ ions react with $Fe^{3+}$ ions to form $FeF_6^{3-}$ ions.
- In Test D, addition of $AgNO_3(aq)$ decreases the amount of $SCN^-$ because $Ag^+$ ions react with $SCN^-$ ions to form a white precipitate of AgSCN(s).
- In Test E, addition of water has no effect.

**MATERIALS**

- 25 mL of $5 \times 10^{-4}$ M iron(III) thiocyanate solution, $Fe(SCN)^{2+}(aq)$*
- 0.1 M iron(III) nitrate solution, $Fe(NO_3)_3(aq)$
- 0.1 M potassium thiocyanate solution, KSCN(aq)
- 0.1 M sodium fluoride solution, NaF(aq)
- 0.1 M silver nitrate solution, $AgNO_3(aq)$
- six semi-micro test tubes
- semi-micro test-tube rack
- bench mat
- marking pen
- dropping pipette
- white tile or white sheet of paper
- safety glasses

*20 mL of 0.1 M $Fe(NO_3)_3(aq)$ and 20 mL of 0.1 M KSCN(aq) per litre

## RISK MANAGEMENT

As directed by your teacher, complete the pre-lab safety information by referring to the safety data sheets (SDSs) or your teacher's risk assessment for the activity.

| Material used | Hazard | Control |
|---|---|---|
| 0.1 M iron(III) nitrate solution, $Fe(NO_3)_3(aq)$ | | |
| 0.1 M sodium fluoride solution, NaF(aq) | | |
| 0.1 M silver nitrate solution, $AgNO_3(aq)$ | can stain skin, clothing and bench surfaces | |

Complete and indicate that you have understood the information in the safety table.

Name (print): _____

I understand the safety information (signature): _____

## METHOD

**1** Fill each of six semi-micro test tubes to one-third of its volume with $Fe(SCN)^{2+}$ solution. Check that the liquid in each tube has the same intensity of colour when you look down the tube, using a white tile or sheet of paper as a background. If necessary, add more solution so that the liquid in each tube is the same colour. Label the tubes A–F.

**2** Using test tube F for the purposes of comparison, perform each of the tests described in Table 1 and record the change that occurs in the colour of the solution when viewed down the test tube.

## VARIABLES

**i** Identify the independent variable: _____

**ii** Identify the dependent variable: _____

**iii** Identify any controlled variables: _____

## RESULTS

**TABLE 1** Colour results

| Test tube | Test | Colour change |
|---|---|---|
| A | 1 drop of $Fe(NO_3)_3$(aq) added | |
| B | 1 drop of KSCN(aq) added | |
| C | 1 drop of NaF(aq) added | |
| D | 1 drop of $AgNO_3$(aq) added | |
| E | equal volume of water added | |
| F | none | no change |



## ANALYSIS

**1** Write an expression for the equilibrium constant of the reaction that is the subject of this practical activity.

_____

**2** Complete Table 2 for each test by stating the:

- **a** initial effect on the concentration of $Fe^{3+}$ or $SCN^-$ of each test
- **b** concentration change of $Fe(SCN)^{2+}$ after the test
- **c** direction in which the position of equilibrium has shifted. The first part has been completed for you.

**TABLE 2** Equilibrium results

| Test tube | Initial effect | $[Fe(SCN)^{2+}]$ change | Direction of equilibrium shift |
|---|---|---|---|
| A | $[Fe^{3+}]$: increases | increases | forward |
| B | $[SCN^-]$: | | |
| C | $[Fe^{3+}]$: | | |
| D | $[SCN^-]$: | | |
| E | $[Fe^{3+}]$: | | |

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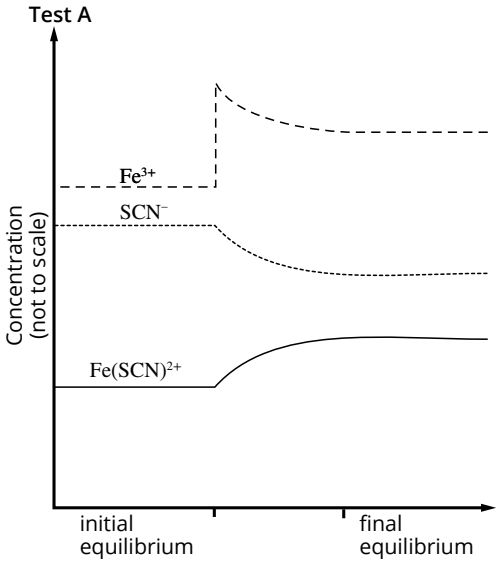
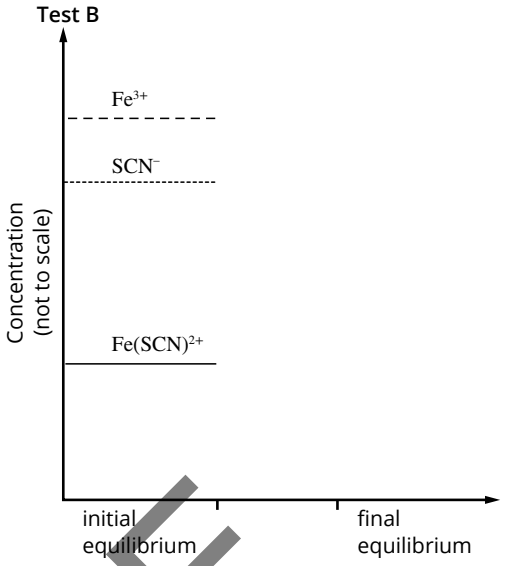
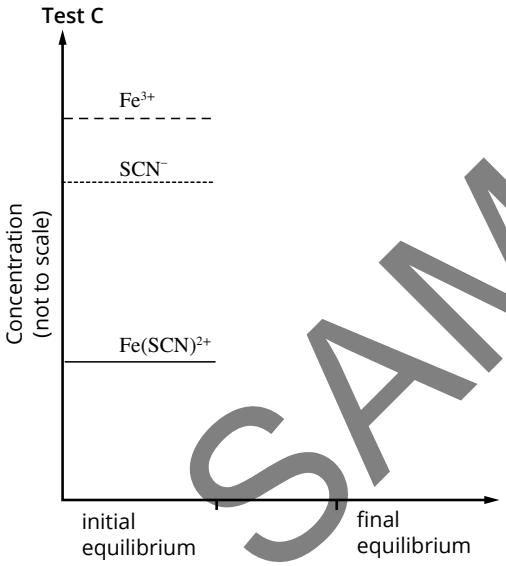
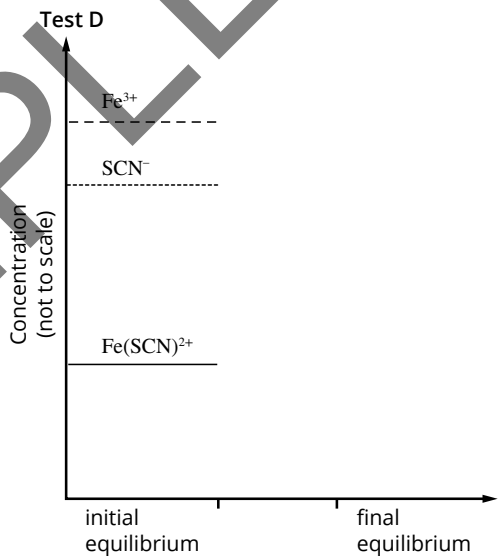
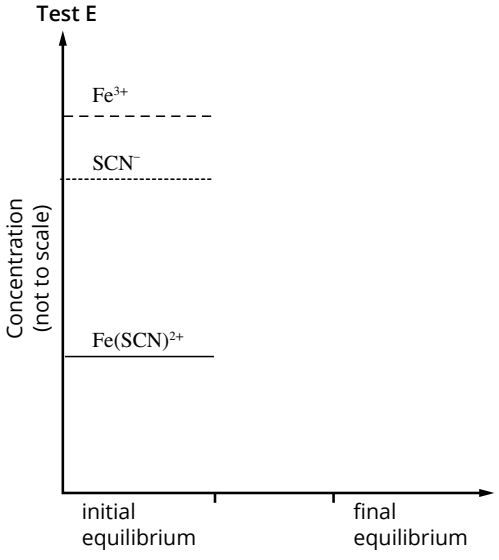
**3** Sketch graphs to show how the concentration of each ion changed during each test. (The graph for Test A has been completed for you.)



**4** Use Le Châtelier's principle to account for the way in which the position of equilibrium shifts in Test A.

**5** Account for the way in which the position of equilibrium shifts in Test E.

**6** Does this method enable these results to be reproducible? Explain briefly how Le Châtelier's principle allows you to be confident that your results are accurate.

## EVALUATION

## CONCLUSION

Summarise the effects of changes in concentration on an equilibrium system.

# EXAM PRACTICE 3.1 • CHEMICAL EQUILIBRIUM SYSTEMS

You may refer to your *QCAA Formula and data book* to answer some questions.

## Multiple-choice questions

**QUESTION 1** *QCAA Chemistry 2023 Paper 1 (Multiple Choice) 1*

In a chemical equation at equilibrium, a reversible arrow (⇌) symbolises that

(A) the forward reaction has stopped, but can be reversed.

(B) the moles of reactants and products present are equal.

(C) half of the reactants have been converted into products.

(D) the concentration of reactants and products remains constant.

**QUESTION 2** *QCAA Chemistry 2023 Paper 1 (Multiple Choice) 2*

Determine which expression represents the hydrogen ion ($H^+$) concentration at a pH of 8.4.

(A) $1 \times 10^{-8.4}$

(B) $1 \times 10^{-5.6}$

(C) $1 \times 10^{-0.9}$

(D) $1 \times 10^{-0.8}$

**QUESTION 3** *QCAA Chemistry 2023 Paper 1 (Multiple Choice) 3*

Two 0.1 M acidic solutions, X and Y, are 100% dissociated. Solution X has an electrical conductivity approximately twice that of solution Y. Identify solutions X and Y.

| | Solution X | Solution Y |
|---|---|---|
| (A) | HCl | $CH_3COOH$ |
| (B) | $HNO_3$ | $H_2SO_4$ |
| (C) | $H_3PO_4$ | $HNO_3$ |
| (D) | $H_2SO_4$ | HCl |

**QUESTION 4** *QCAA Chemistry 2022 Paper 1 (Multiple Choice) 11*

A Brønsted–Lowry acid

(A) accepts a proton to form its base.

(B) donates a proton to form its base.

(C) accepts a proton to form its conjugate base.

(D) donates a proton to form its conjugate base.

**QUESTION 5** *QCAA Chemistry 2021 Paper 1 (Multiple Choice) 10*

Determine which system at equilibrium will shift to the right (products) if the total pressure on the system is increased.

(A) $N_2O_4(g) \rightleftharpoons 2NO_2(g)$

(B) $2H_2O(g) \rightleftharpoons 2H_2(g) + O_2(g)$

(C) $2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$

(D) $CO(g) + H_2O(g) \rightleftharpoons CO_2(g) + H_2(g)$



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**QUESTION 6** *QCAA Chemistry 2024 Paper 1 (Multiple Choice) 10*

The acid dissociation constant ($K_a$) represents the

(A) pH of an acid solution.

(B) strength of an acid solution.

(C) concentration of an acid solution.

(D) conjugate acid–base pairs of an acid solution.

**QUESTION 7** *QCAA Chemistry 2024 Paper 1 (Multiple Choice) 5*

Titration is a volumetric analysis method used to

(A) measure the pH of an analyte.

(B) determine the volume of a titrant.

(C) calculate the concentration of an identified solution.

(D) prepare a standard solution of known concentration and volume.

**QUESTION 8** *QCAA Chemistry 2024 Paper 1 (Multiple Choice) 6*

The equivalence point of an acid–base titration occurs when the

(A) pH equals the p$K_a$.

(B) pH stops changing.

(C) indicator changes colour.

(D) titrant completely neutralises the analyte.

**QUESTION 9** *QCAA Chemistry 2022 Paper 1 (Multiple Choice) 10*

The midpoint of the colour change of a weak acid indicator occurs when

(A) $[In^-] = [H^+]$

(B) $[In^-] = [HIn]$

(C) $[H^+] = [OH^-]$

(D) $[HIn] = [OH^-]$

**QUESTION 10** *QCAA Chemistry 2022 Paper 1 (Multiple Choice) 8*

Predict how the system shown will respond when a small amount of aqueous sodium hydroxide is added.

$$CH_3COOH(aq) + H_2O(l) \rightleftharpoons CH_3COO^-(aq) + H_3O^+(aq)$$

(A) Equilibrium shifts to the left, and the pH decreases.

(B) Equilibrium shifts to the right, and the pH increases.

(C) Equilibrium shifts to the left, and the pH remains the same.

(D) Equilibrium shifts to the right, and the pH remains the same.

## Short-response questions

**QUESTION 1 (3 marks)** *QCAA Chemistry 2024 Paper 2 (Short Response) 1*

A 30.00 mL aliquot of aqueous hydrochloric acid (HCl) with a pH of 2.00 was diluted with 2970.00 mL of water to make up a final volume of 3.00 L of aqueous HCl solution.

a) Determine the concentration of hydrogen ions in the original 30.00 mL aliquot of HCl(aq). *[1 mark]*

b) Calculate the pH of the final 3.00 L solution of HCl(aq). Show your working. *[2 marks]*

**QUESTION 2 (6 marks)** *QCAA Chemistry 2024 Paper 2 (Short Response) 2*

Phenol red (HIn) is a weak acid that acts as an indicator, as shown.

$$HIn(aq) \rightleftharpoons In^-(aq) + H^+(aq)$$

a) Identify the conjugate base of phenol red. _____ *[1 mark]*

b) Determine the dissociation constant ($K_a$) of phenol red. *[1 mark]*

c) Explain the relationship between the pH range of colour change for phenol red and its p$K_a$ value. *[4 marks]*

**QUESTION 3 (5 marks)** *QCAA Chemistry 2021 Paper 1 (Short Response) 25*

An equilibrium is formed between two differently coloured cobalt species, $Co(H_2O)_6^{2+}(aq)$, which is pink, and $CoCl_4^{2-}(aq)$, which is blue. The equation for this equilibrium is shown.

$$Co(H_2O)_6^{2+}(aq) + 4Cl^-(aq) \rightleftharpoons CoCl_4^{2-}(aq) + 6H_2O(l)$$

a) Apply Le Châtelier's principle to predict the visible effect of adding $AgNO_3$ to an aqueous blue-coloured solution containing $Co(H_2O)_6^{2+}$ and $CoCl_4^{2-}$ ions. Explain your reasoning. *[3 marks]*

b) When a sample of the equilibrium mixture is put into hot water, the mixture turns more blue. Determine whether the forward reaction of the equation is exothermic or endothermic. Explain your reasoning. *[2 marks]*

**QUESTION 4 (13 marks)** *QCAA Chemistry 2023 Paper 2 (Short Response) 5*

The table gives the properties of four monoprotic acids.

| Acid | Concentration (mol $L^{-1}$) | $[H^+]$ (mol $L^{-1}$) | pH | $K_a$ |
|---|---|---|---|---|
| 1 | 0.200 | $7.90 \times 10^{-5}$ | | |
| 2 | 0.100 | $4.20 \times 10^{-3}$ | 2.34 | $1.80 \times 10^{-4}$ |
| $CH_3COOH(aq)$ | 0.100 | | | $1.78 \times 10^{-5}$ |
| HCl(aq) | 0.010 | $1.00 \times 10^{-2}$ | 2.00 | >1 |

a) Determine the relative strength of acids 1 and 2 by contrasting their $K_a$ values. *[3 marks]*

b) Write a balanced chemical equation for the dissociation of ethanoic acid ($CH_3COOH$) in water. *[2 marks]*

c) Identify whether the conjugate base of ethanoic acid ($CH_3COOH(aq)$) is amphiprotic. Explain your reasoning. *[2 marks]*

d) Calculate the pH of the aqueous solution of ethanoic acid ($CH_3COOH$). Show your working. *[3 marks]*

e) Determine the volume of water that would need to be added to 100.0 mL of HCl(aq) to change the pH from 2.00 to 3.00. Explain your reasoning. *[3 marks]*

**QUESTION 5 (2 marks)** *QCAA Chemistry 2024 Paper 2 (Short Response) 8*

In the aqueous solution of a chromate salt, an equilibrium exists between the yellow chromate ions and the orange dichromate ions. This equilibrium can be represented by the equation shown.

$$2CrO_4^{2-}(aq) + 2H^+(aq) \rightleftharpoons Cr_2O_7^{2-}(aq) + 2H_2O(l)$$

yellow orange

Explain, at an atomic level, why no colour change occurs once the chromate–dichromate solution has established equilibrium.

# SAMPLE ASSESSMENT TASK IA1

# Data test (60 marks)

**Suggested duration:** 10 minutes reading time and 60 minutes to complete the test

## Task

The data test requires you to apply a range of cognitions (ways of gaining knowledge and understanding) to respond to scientific data. The test will be held in a set time frame under test conditions.

Each data set will enable you to analyse and interpret data to apply your understanding of equilibrium and redox reactions.

You will be required to complete short- and long-answer questions, calculations and interpretation of graphs to a total of up to 400–500 words. Complete all parts of each question as directed.

## Data set 1: Investigating the equilibrium law (10 marks)

Chemists investigating the equilibrium law obtained the data in Table 1 for the gaseous equilibrium between nitrogen tetroxide and nitrogen dioxide:

$$N_2O_4(g) \rightleftharpoons 2NO_2(g)$$

Different starting concentrations of gaseous $N_2O_4$ and $NO_2$ were used in four separate experiments, so different equilibrium concentrations were obtained. The experiments were conducted in a sealed vessel at 25°C.

**TABLE 1** Concentrations of reactants and products present in equilibrium mixtures

| Equilibrium mixture | $[N_2O_4]$ in mol $L^{-1}$ | $[NO_2]$ in mol $L^{-1}$ |
|---|---|---|
| 1 | 0.0016 | 0.0010 |
| 2 | 0.0155 | 0.0031 |
| 3 | 0.0570 | 0.0060 |
| 4 | 0.1950 | 0.0110 |



**Item 1:** From your knowledge of chemical equilibrium, show the expression for the equilibrium law for this reaction in terms of the concentrations of the reactants and products. (1 mark)

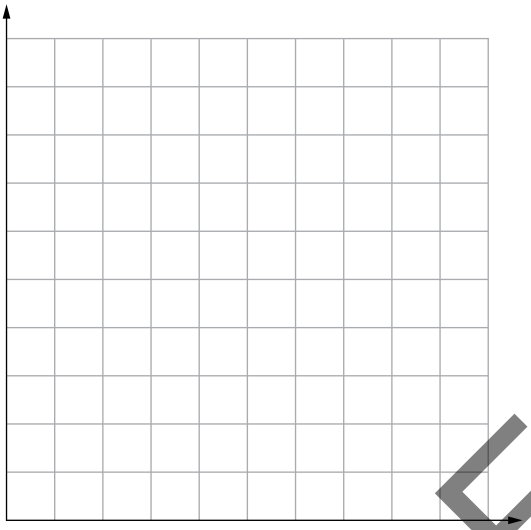
**Item 2:** Use the data to calculate the value of the equilibrium constant for each mixture, then determine the average value of the equilibrium constant at 25°C. (2 marks)

**Item 3:** Calculate the equilibrium constant at 25°C for the reaction:

$$4NO_2(g) \rightleftharpoons 2N_2O_4(g)$$ (2 marks)

**Item 4:** Construct a concentration–time graph for the equilibrium system in **Item 1** to show the effects if more nitrogen dioxide gas is pumped into the system five minutes after the system has achieved equilibrium. Label your graph using the words and formulas below. (5 marks)

| Concentration | Time | $NO_2$ added | $[NO_2]$ | $[N_2O_4]$ | initial equilibrium | new equilibrium established |
|---|---|---|---|---|---|---|



## Data set 2: Understanding galvanic cells and the use of the table of standard electrode potentials (15 marks)

Use the table of standard electrode potentials in your data booklet to answer these questions.

**Item 1:** Predict the anode and cathode reactions when a cell is constructed using $Mg^{2+}(aq)/Mg(s)$ and $Cu^{2+}(aq)/Cu(s)$ half-cells at standard conditions. Show the half-equations below. (2 marks)

Anode: _____

Cathode: _____

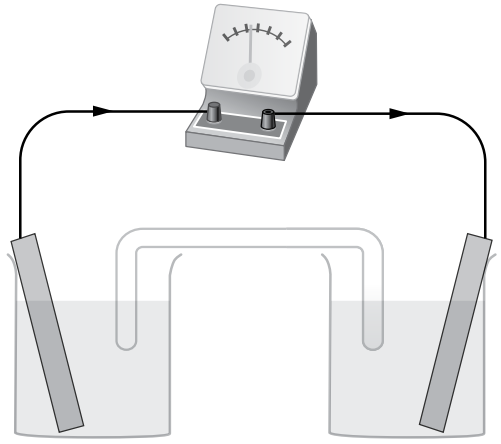
**Item 2:** Determine the equation for the overall cell reaction and calculate the cell potential. (2 marks)

Overall reaction: _____

Cell potential: _____

**Item 3:** Label the diagram using the words below. (3 marks)

| | | | | | |
|---|---|---|---|---|---|
| anode | cathode | copper electrode | $Mg^{2+}$ ions | positive electrode | electrode where oxidation occurs |
| salt bridge | $Cu^{2+}$ ions | negative electrode | magnesium electrode | electrode where reduction occurs | |



**Item 4:** On the diagram for **Item 3**, indicate the direction of movement of electrons in the external circuit and cations in the salt bridge. (2 marks)

**Item 5:** Determine whether reactions occur in **a–d** below and show overall redox equations for those reactions that do occur. (6 marks)

**a** A piece of iron is placed in a solution containing $Sn^{2+}$ ions.

**b** A piece of lead is placed in a solution containing $Fe^{2+}$ ions.

**c** A piece of zinc metal is placed in a solution of $Mg^{2+}$ ions.

**d** A piece of zinc is placed in a beaker of oxygenated water.

## Data set 3: Using volumetric analysis (9 marks)

A student determined the concentration of ethanoic acid in vinegar by volumetric analysis. The student used a pipette to deliver 20.00 mL of vinegar into a conical flask, and the vinegar was then titrated with 1.00 mol $L^{-1}$ sodium hydroxide solution. This was repeated three times. The student's results are shown in Table 2.

**TABLE 2** Titration results for ethanoic acid against sodium hydroxide solution

| Titration number | Initial burette reading (mL) | Final burette reading (mL) | Titre (mL) of sodium hydroxide solution |
|---|---|---|---|
| 1 | 0.00 | 21.00 | 21.00 |
| 2 | 21.00 | 41.10 | 20.10 |
| 3 | 0.00 | 19.99 | 19.99 |
| 4 | 19.99 | 40.12 | 20.13 |

**Item 1:** Write a balanced equation for the reaction between ethanoic acid and sodium hydroxide solution. (1 mark)

**Item 2:** Identify the concordant titres and use these to calculate the average titre of the sodium hydroxide solution used during the titration. (2 marks)

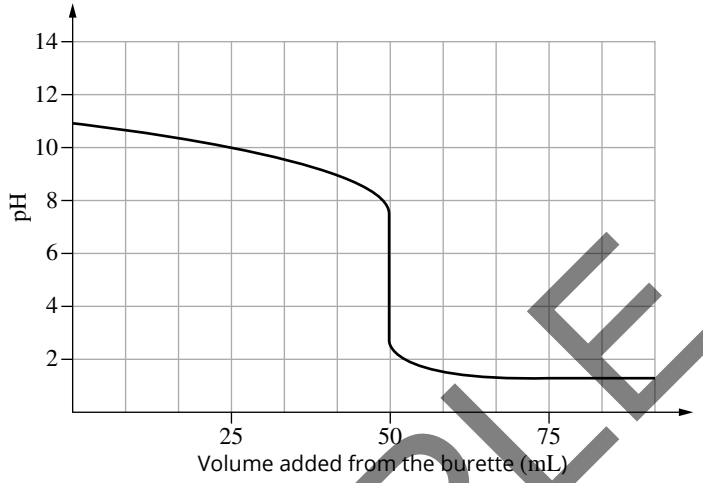
**Item 3:** Calculate the number of moles of sodium hydroxide in the average titre of sodium hydroxide solution. (2 marks)

**Item 4:** Determine the number of moles of ethanoic acid present in each 20.00 mL aliquot of vinegar. (2 marks)

**Item 5:** Calculate the concentration of ethanoic acid (in mol $L^{-1}$) in the sample of vinegar. (2 marks)

## Data set 4: Understanding acid–base titration curves (13 marks)

An acid–base titration curve for a 1 mol $L^{-1}$ base and a strong acid is shown below. Use data from the curve to answer the following questions.




**Item 1:** Is the base is in the conical flask or in the burette? Explain. (1 mark)

**Item 2:** State whether the base is strong or weak. Explain your answer. (2 marks)

**Item 3:** Label the following points or regions on the graph. (4 marks)

- equivalence point
- the point where pH = p$K_a$
- buffer region
- volume of acid used to reach the equivalence point.

**Item 4:** Estimate the $K_b$ of the base from the graph. (2 marks)

**Item 5:** Describe the main characteristics of a buffer solution. Using Le Châtelier's principle, explain how buffer solutions behave when a small amount of acid or base is added. (4 marks)

## Data set 5: Investigating electroplating (13 marks)

Two students were given the chemicals and equipment needed to zinc plate a toy metal car. They constructed an electroplating cell using a large beaker with 60 mL of 1 mol $L^{-1}$ $ZnSO_4$ solution and a zinc strip. Their results and some additional data are given in the following table.

| Some data and results of zinc-plating experiment | |
|---|---|
| Mass of the toy car before plating | 105.0 g |
| Mass of the toy car after plating | 108.3 g |
| Molar mass of Zn | 65.4 g $mol^{-1}$ |

**Item 1:** Write the equation for the reaction occurring at the cathode in the zinc-plating cell. (2 marks)

**Item 2:** Calculate the amount of zinc deposited, in mol, on the toy car. (2 marks)

**Item 3:** Calculate the number of moles of electrons needed to deposit this mass of zinc. (2 marks)

**Item 4:** Calculate the current, in A, that was needed to deposit this mass of zinc. (3 marks)

**Item 5:** The electrical energy consumed by a cell is related to the cell voltage, current and time of operation by the formula $E = VIt$. Calculate the energy, in J, consumed by the cell to deposit this mass of zinc. (4 marks)



| RATE MY LEARNING | • I get it.<br>• I can apply/teach it. | • I get it.<br>• I can show I get it. | • I almost get it.<br>• I might need help. | • I get some of it.<br>• I need help. | • I don't get it.<br>• I need lots of help. |
|---|---|---|---|---|---|

# SAMPLE ASSESSMENT TASK IA2

# Student experiment

**Suggested duration**: up to 10 hours (including development of research question, practical work, writing time and preparation of report)

## Introduction

The student experiment requires you to apply the scientific method over an extended and defined period of time. You will develop your own research question or hypothesis to investigate, based on an initial experiment already completed in class.

The student experiment forms an important component of the assessment in this unit. It is an opportunity to further explore concepts already covered and of particular interest to you, with some latitude for self-direction. The task is estimated to take approximately 10 hours of class time. It is suggested that the time is divided approximately as follows.

- **Research and planning**: 3–4 hours
- **Conducting the experiment**: 2 hours
- **Final report**: 4–5 hours

Discussing your experiment with your teacher and receiving feedback will be an important part of the planning process.

You may develop the method and conduct the experiment collaboratively with another student if you wish. All other stages of your student experiment, such as analysis of data, must be carried out individually.

## Report requirements

You will produce a scientific report for this assessment task with a word limit of 2000 words.

Your report is a formal scientific report that requires you to demonstrate the research and planning you have undertaken in preparing for the experimental investigation. It will be helpful to prepare your report using specific headings. The following plan has been prepared to guide you through the task. This sample student experiment provides a template to assist you in completing the task. Some elements of this sample assessment task have been completed for you.

The report can also be presented as a multimodal presentation (e.g. a scientific poster presentation combined with a speech). Your teacher can provide more advice about this option.

## Task

You have completed a number of practical activities in Unit 3. This assessment task requires you to modify (refine, extend or redirect) a practical activity of your choice to address your own related research question.

Your research will involve collecting and analysing primary data. Undertake a literature survey of background information related to your research question, which will assist you in the analysis and discussion of your results. You may also need to locate and use information beyond your own knowledge and the data you have been given, and use this information in conjunction with your experimental primary data to answer your research question.

The following questions will help guide the design of your investigation. All of your notes must be recorded in your laboratory notebook, but this section should help you visualise and plan the entire process. Copy the following headings into your notebook and use these headings when writing your report.

## Research and planning

**Research question**—Develop a specific question that will form the basis for planning the investigation. The question must be relevant to and drawn from the theme of the original experiment, and must include specific and measurable variables. The steps below will help you in developing your research question.

## SAMPLE ASSESSMENT TASK IA2

**1** Name the original experiment that you will modify.

**2** Select one factor, or an extension of the activity, which you would like to investigate further. You may find it helpful to refer to the aim of the experiment that you completed previously in class.

**3** Briefly explain the theory behind the modifications you are making to the procedure of the original class practical activity.

**4** **a** State your draft research question clearly, including the independent and dependent variables and the variables to be kept constant during the experiment.

| Independent variable and how it is changed | Dependent variable and how it is measured | Controlled variables |
|---|---|---|
| | | |

**b** Reflect on your draft research question. Can the wording be refined and focused to better guide your experiment? Show your refined research question to your teacher.

**Rationale**—Introduce the investigation and explain the conceptual understanding that led to the development of the research question. The steps below will help you.

**5** Identify the scientific terms and concepts that are relevant to the original experiment and that are also relevant to your research question.

**6** Indicate the references you will use to show the extent of your literature survey and research.

**7** Summarise the modifications to the method of the original experiment that you are making. Provide an explanation of the theory relevant only to the modifications.

**Method**—Explain in detail how you plan to modify the method of the original experiment to meet the new line of inquiry.

**8** For your planning, list the additional materials and equipment you will need.

**9** Outline the method to be used for your experiment, which will refine and extend the original experiment. When describing the procedure for your investigation:

- show the data needed so you can draw suitable conclusions and answer your question fully
- if appropriate, state the number of trials required to reduce random errors and enable reliable trends and relationships to be analysed.

**Risk assessment**—Detail how you will address safety hazards and ethical considerations of the modifications to minimise the risks.

**10** Consider the risks associated with the chemicals and the equipment. Show your risk management program as well as procedures for managing any adverse events that may occur.

## Results and analysis

### EXPERIMENTAL DATA

When conducting the experiment, make sure you collect sufficient qualitative and quantitative information to enable you to answer your question accurately.

**11** Under the heading 'Results', show your results in appropriate tables with informative headings and units as needed.

**12** Under the heading 'Experimental arrangement', draw diagrams of the equipment you used, and clearly label them using scientific notation and terms. You could also use labelled photographs.



### PROCESSING AND ANALYSING

**13** Use graphs, tables and calculations (if necessary) to present the information and results of your analysis. Where appropriate, use the correct significant figures and include units. Identify the best ways to analyse the raw and processed data to enable any trends, patterns or relationships to be evident.

**14** Use the data to identify key trends, patterns or relationships.

**15** Identify any limitations to the trends, patterns or relationships you have identified, including any random and systematic errors present in the data.

## Interpreting and evaluation

### CONCLUSION

**16** Use the results to answer your research question. State the evidence from your results that allows you to answer your question. List the reasons for any relationships you have observed. The trends, patterns or relationships you have identified from the data should be related to the chemical concepts of the experiment.

### EVALUATION AND SOURCES OF ERROR

**17** Summarise the limitations of the evidence and data. Identify possible errors and suggest why they may have occurred. Evaluate the reliability and validity of your experimental process to determine if:

- your method is consistent and repeatable and provides possible reason(s) for why errors occurred
- you have obtained applicable and defensible results to answer the question.

### SUGGESTED IMPROVEMENTS AND EXTENSIONS

**18** Identify possible improvements and extensions for future experiments for each of the sources of errors you described above. Make reference to your literature research.

### REFERENCE LIST

**19** Use references throughout the report when quoting data or evidence. Such references can include referring to data and comments in your laboratory notebook where you recorded information and results. Provide a list of references at the end of your report.

## Communication—presenting your findings

Writing the scientific report is an important final part of this investigation.

- Carefully plan your headings and presentation of data and results, as well as your conclusions, in conjunction with the requirements of the task.
- Make sure that the correct significant figures and units have been used.
- Including a word count is a useful way to indicate your consideration of the report-writing requirements.

The following subheadings will give you an idea of what you should include in your scientific report.

### RESEARCH AND PLANNING

#### Research question

- Is your research question clear and concise?

#### Rationale

- Have you identified and defined relevant background concepts with in-text citations to appropriate references?
- Have you described the original experiment and made links to the development of your own research question?
- Have you identified and justified the independent, dependent and controlled variables?

#### Method

- Have you described and justified all modifications you have made to the original method?
- Have you comprehensively selected new equipment and materials?
- Does your method comprehensively meet your aim?
- Is your method repeatable?
- Where possible, have you included a labelled diagram of experimental set-up/results?
- Have you labelled each figure (and included a reference where applicable)?

#### Risk assessment

- Have you identified, explained and managed any risks?
- Have you incorporated safe and ethical work practices?

### RESULTS AND ANALYSIS

#### Raw data

- Have you collected qualitative observations as well as raw data?
- Have you collected relevant data systematically?
- Have you collected sufficient and relevant data?
- Is your raw data clear, accurate and detailed?
- Have you processed your data in the most appropriate ways to illustrate trends, patterns and/or scientific relationships?

#### Processed data

- Have you comprehensively explained how accuracy, precision, reliability, validity, uncertainty and errors relate to your collected data?
- Have you analysed and explained your investigation data?

**Analysis**

- Does the discussion of your analysis address all possible sources of error and uncertainty?
- Have you considered how your analysis may be limited by the data you did or did not collect?

**INTERPRETING AND EVALUATION**

**Conclusion**

- Is your conclusion valid given your analysis and evaluation?
- Does your conclusion relate collected data to the hypothesis and/or aim?
- Does it provide a justified response to the research question?
- Have you explained the link between your investigation findings and the relevant scientific concepts you included in the Research and planning section?
- Have you used scientific language and conventions?

**Evaluation and sources of error**

- Were the procedures and equipment you used appropriate and effective?

**Improvements and extensions**

- Have you made practical suggestions for improvements to the experimental design?
- Have you suggested possible further investigations?

**Reference list**

- Have you used a standard scientific referencing system?
- Is the terminology and structure of the report written in a scientific manner?



| RATE MY LEARNING | • I get it.<br>• I can apply/teach it. | • I get it.<br>• I can show I get it. | • I almost get it.<br>• I might need help. | • I get some of it.<br>• I need help. | • I don't get it.<br>• I need lots of help. |
|---|---|---|---|---|---|