

PEARSON
CHEMISTRY

NEW SOUTH WALES

SKILLS AND ASSESSMENT



Penny Commons

NSW
STAGE 6

Contents

Chemistry toolkit

Module 5: Equilibrium and acid reactions

KEY KNOWLEDGE	3
WORKSHEETS	
WORKSHEET 5.1 Knowledge review—thinking about rates and energy	12
WORKSHEET 5.2 Exploring equilibrium—reactions in the balance	13
WORKSHEET 5.3 Calculations—equilibrium constants and concentrations	16
WORKSHEET 5.4 Equilibrium—Le Châtelier’s principle and the equilibrium law	18
WORKSHEET 5.5 Wonderful water—structure and properties	20
WORKSHEET 5.6 Solving solubility—predicting precipitation reactions and writing ionic equations	21
WORKSHEET 5.7 Solubility equilibrium—calculating solubility and K_{sp}	22
WORKSHEET 5.8 Predicting precipitation using K_{sp}	23
WORKSHEET 5.9 Literacy review—equilibrium terms and expressions	24
WORKSHEET 5.10 Thinking about my learning	25
PRACTICAL ACTIVITIES	
ACTIVITY 5.1 Reversible and irreversible reactions	26
ACTIVITY 5.2 Investigation of the cobalt chloride equilibrium system	30
ACTIVITY 5.3 Effect of temperature on equilibrium yields	32
ACTIVITY 5.4 Determination of K_{eq} by colorimetry	34
ACTIVITY 5.5 Precipitation reactions	37
DEPTH STUDY 5.1 Practical investigation of an equilibrium system	41
MODULE 5 REVIEW QUESTIONS	46

ix Module 6: Acid/base reactions

KEY KNOWLEDGE	50
WORKSHEETS	
WORKSHEET 6.1 Knowledge review—acids and bases	61
WORKSHEET 6.2 Concentration and strength—modelling acids and bases and calculating pH	63
WORKSHEET 6.3 Acid–base relationships—calculating pH and pOH	65
WORKSHEET 6.4 Indicators and K_a —colourful considerations	66
WORKSHEET 6.5 Calculating the pH of mixtures and dilutions	67
WORKSHEET 6.6 Acid–base curves—detecting an equivalence point	68
WORKSHEET 6.7 Volumetric analysis—finding the fizz in soda water	70
WORKSHEET 6.8 Buffers—keeping pH in control	72
WORKSHEET 6.9 Literacy review—acids, bases and volumetric analysis	73
WORKSHEET 6.10 Thinking about my learning	74
PRACTICAL ACTIVITIES	
ACTIVITY 6.1 Comparing the strengths of acids and bases	75
ACTIVITY 6.2 Amphiprotic substances in water	78
ACTIVITY 6.3 Reactions of hydrochloric acid	80
ACTIVITY 6.4 Measuring the enthalpy of neutralisation	83
ACTIVITY 6.5 Beetroot—a natural indicator	85
ACTIVITY 6.6 Determination of the ethanoic acid concentration of vinegar	88
ACTIVITY 6.7 Determination of two acid dissociation constants	91
ACTIVITY 6.8 Buffers	93
DEPTH STUDY 6.1 Practical investigation—focus on acids and bases	95
MODULE 6 REVIEW QUESTIONS	97

Contents

Module 7: Organic chemistry

KEY KNOWLEDGE 102

WORKSHEETS

WORKSHEET 7.1	Knowledge review—molecular shape and intermolecular forces	118
WORKSHEET 7.2	Families of hydrocarbons—alkanes, alkenes and alkynes	120
WORKSHEET 7.3	Families examined—properties and structure	123
WORKSHEET 7.4	Observing organics—properties of functional groups	125
WORKSHEET 7.5	Converting chemicals—organic reaction pathways	126
WORKSHEET 7.6	Focus on functional groups—effect on chemical reactions	128
WORKSHEET 7.7	Biofuels—biogas, bioethanol and biodiesel	131
WORKSHEET 7.8	Consumer products—polymers and soaps	134
WORKSHEET 7.9	Literacy review—reviewing organic chemistry	137
WORKSHEET 7.10	Thinking about my learning	139

PRACTICAL ACTIVITIES

ACTIVITY 7.1	Investigating hydrocarbons	140
ACTIVITY 7.2	Modelling hydrocarbons, functional groups and organic reactions	143
ACTIVITY 7.3	Determining the heat of combustion of alcohols	147
ACTIVITY 7.4	Preparing artificial fragrances and flavours	150
ACTIVITY 7.5	Modelling polymers	152
DEPTH STUDY 7.1	What is the importance of organic chemistry to society?	154

MODULE 7 REVIEW QUESTIONS 156

Module 8: Applying chemical ideas

KEY KNOWLEDGE 160

WORKSHEETS

WORKSHEET 8.1	Knowledge review—atomic structure and stoichiometry	170
WORKSHEET 8.2	Puzzling about ions—testing techniques	171
WORKSHEET 8.3	Colorimetry and UV-visible spectroscopy	172
WORKSHEET 8.4	Atomic absorption spectroscopy	174
WORKSHEET 8.5	Forensic testing—blood alcohol concentration	175
WORKSHEET 8.6	Nuclear magnetic resonance	177
WORKSHEET 8.7	Solving a chemical mystery using spectroscopy	179
WORKSHEET 8.8	Optimising reactions in industry	182
WORKSHEET 8.9	Literacy review—understanding chemical ideas	184
WORKSHEET 8.10	Thinking about my learning	187

PRACTICAL ACTIVITIES

ACTIVITY 8.1	Testing for ions in aqueous solutions	188
ACTIVITY 8.2	Gravimetric determination of sulfur as sulfate in fertiliser	192
ACTIVITY 8.3	Analysing precipitation titration data	195
ACTIVITY 8.4	Colorimetric analysis of iron in foods	197
ACTIVITY 8.5	Analysing UV-visible spectroscopy data	200
ACTIVITY 8.6	Analysing atomic absorption spectroscopy data	202
ACTIVITY 8.7	Testing for functional groups	204

DEPTH STUDY 8.1	Investigating the applications of chemical ideas	207
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MODULE 8 REVIEW QUESTIONS 210

PERIODIC TABLE 214

How to use this book

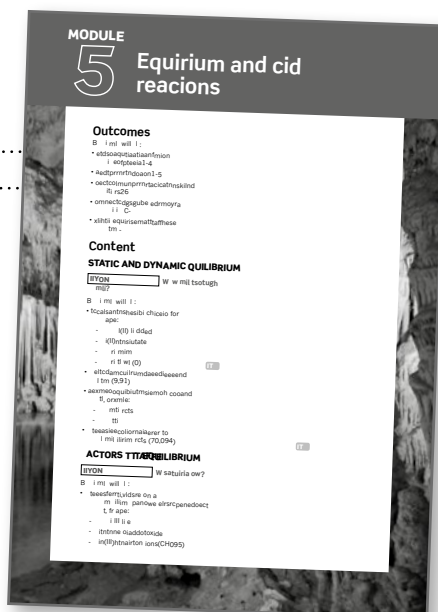
The *Pearson Chemistry 12 New South Wales Skills and Assessment* book provides an intuitive, self-paced approach to science education that ensures every student has opportunities to practise, apply and extend their learning through a range of supportive and challenging activities. While offering opportunities for reinforcement of key concepts, knowledge and skills, these activities enable flexibility in the approach to teaching and learning.

Explicit scaffolding makes learning objectives clear, and there are regular opportunities for student reflection and self-evaluation at the end of individual activities throughout the book. Students are also guided in self-reflection at the end of each module. In addition, there are rich opportunities to take the content further with the explicit coverage of Working scientifically skills and key knowledge in the depth studies.

This resource has been written to the new Stage 6 Syllabus for New South Wales Chemistry and addresses the final four modules of the syllabus. Each module consists of five main sections:

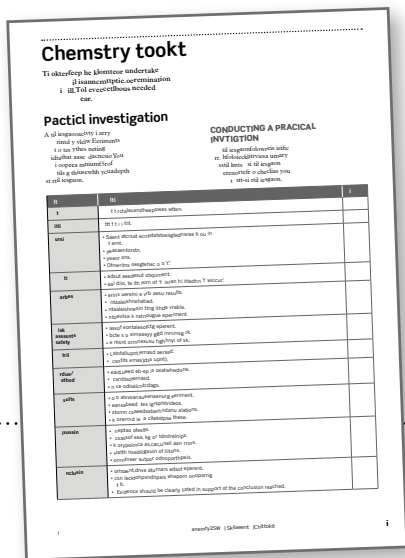
- key knowledge
- worksheets
- practical activities
- depth study
- module review questions.

Explore how to use this book below.



Chemistry toolkit

The Chemistry toolkit supports development of the skills and techniques needed to undertake practical investigations secondary-sourced investigations and depth studies and covers examination techniques and study skills. It also includes checklists, exemplars and scaffolded steps. The toolkit can serve as a reference tool to be consulted as needed.

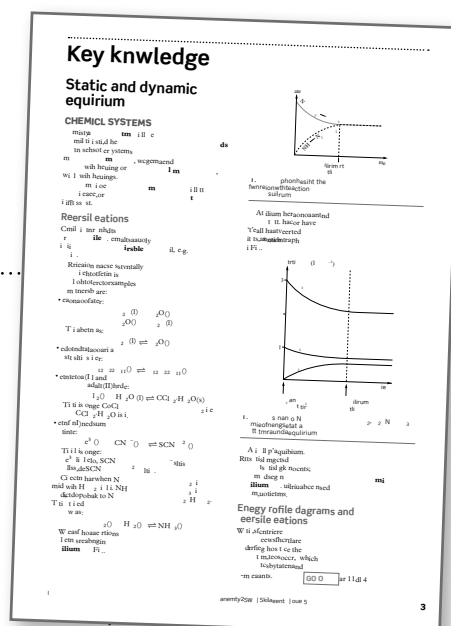


Module opener

Each book is split into the four modules of the syllabus with the module opener linking the module content to the syllabus.

Key knowledge

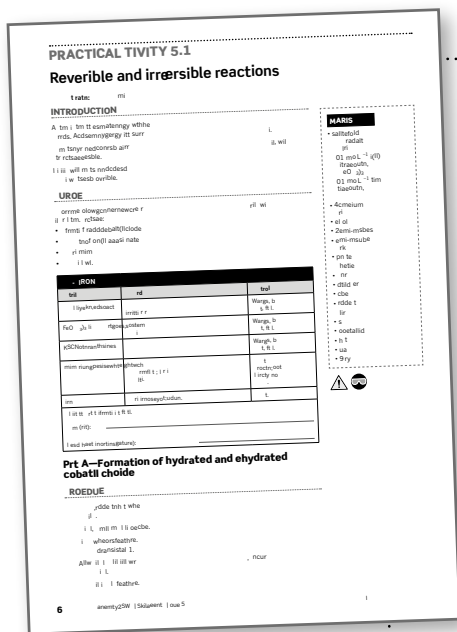
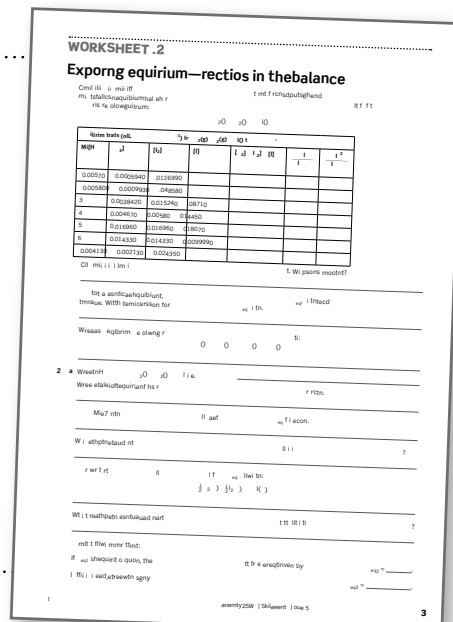
Each module begins with a key knowledge section. This consists of a set of succinct summary notes that cover the key knowledge set out in each module of the syllabus. This section is highly structured and written in a straightforward style to assist students of all reading abilities. Key terms are bolded for ease of navigation. It also provides a ready reference for completing the worksheets and practical activities.



Worksheets

A diverse offering of instructive and self-contained worksheets is included in each module. Common to all modules is the 'Knowledge review' worksheet to activate prior knowledge, a 'Literacy review' worksheet to explicitly build understanding and application of scientific terminology, and finally a 'Thinking about my learning' worksheet which provides a reflection and self-assessment opportunity for students. Each additional worksheet provides opportunities to review, consolidate and further student understanding.

These worksheets function as formative assessment and are clearly aligned to the syllabus. A range of questions building from foundation to challenging are included within worksheets.



Practical activities

Practical activities provide the opportunity to complete practical work related to the various themes covered in the syllabus. All practical activities referenced in outcomes within the syllabus have been covered. Across the suite of practical activities provided, students are exposed to opportunities where they design, conduct, evaluate, gather and analyse data appropriately record results and prepare evidence-based conclusions directly into the scaffolded practical activities. Students also have opportunities to evaluate safety and any potential hazard.

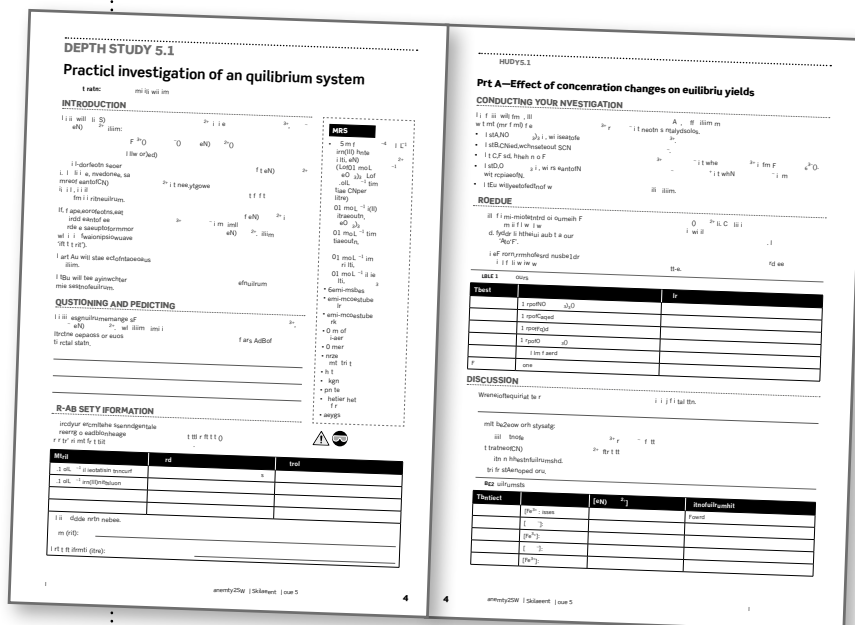
Each practical activity includes a suggested duration. Along with the depth studies the practical activities meet the 35 hours of practical work mandated at Year 12 in the syllabus. Where there is key knowledge that will support the completion of a practical activity, students are referred back to it.

Like the worksheet, the practical activities include a range of questions building from foundation to challenging.

Depth studies

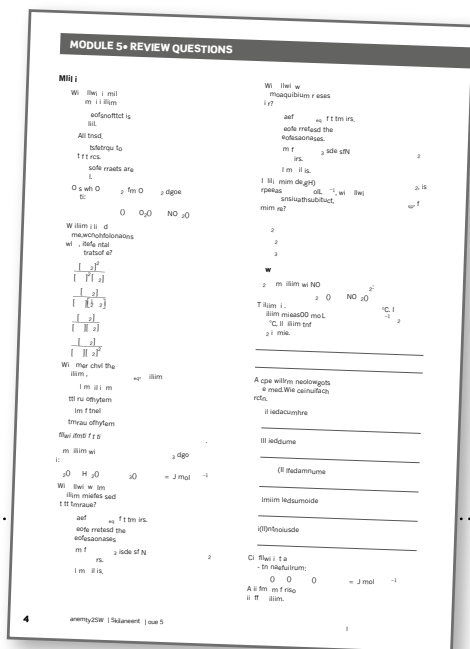
Each module contains one suggested depth study. The depth studies allow further development of one or more concepts found within or inspired by the syllabus. They allow students to acquire a depth of understanding and take responsibility for their own learning. They also promote differentiation and engagement.

Each depth study allows for the demonstration of a range of Working Scientific Skills with a depth studies assessing the Working Scientifically outcomes of Questioning and Predicting, and Communicating. A minimum of two additional Working Scientifically skills and at least one Knowledge and Understanding outcome are also assessed.



Module review questions

Each module finishes with a comprehensive set of questions consisting of multiple choice and short answer, that helps students to draw together their knowledge and understanding and apply it to these styles of questions



Icons and features

The New South Wales Stage 6 Syllabus Learning across the curriculum content is addressed and identified using the following icons



GoTo icons are used to make important links to relevant content within the book



The **safety con** highlights significant hazards indicating caution is needed.



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

A **prelab safety box** is included. Students are to sign agreeing that they have understood the hazards associated with the material(s) in use and the control measures to be taken.

PRE-LAB SAFETY INFORMATION

Material	Hazard	Control
KSCN solution	irritating to the skin and eyes	Wear gloves, lab coat, safety glasses.
FeSCN ²⁺ solution	contains nitric acid and is irritating to the skin and eyes	Wear gloves, lab coat, safety glasses.
F(NO ₃) ₃ solution	irritating to eyes, respiratory system and skin	Wear gloves, lab coat and safety glasses.

Please indicate that you have understood the information in the safety table.

Name (print): _____

I understand the safety information (signature): _____

Rating my learning

Rating my learning is an innovative tool that appears at the bottom of the final page of most worksheets and all practical activities. It provides students with the opportunity for self-reflection and self-assessment. It encourages students to look ahead to how they can continue to improve, and assists in highlighting focus areas for further skill and knowledge development.

The teacher may choose to use student responses to the 'Rating my learning' feature as a formative assessment tool. At a glance, teachers can assess which topics and which students need intervention for improvement.

RAISING MY LEARNING	My understanding improved	No confident	Vry confident	answered questions without help	No confident	Vry confident	corrected my errors without help	No confident	Vry confident
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Teacher support

Comprehensive answers and fully worked solutions for all worksheets, practical activities, depth studies and module review questions are provided in the *Pearson Chemistry 12 New South Wales Teacher Support*. An editable suggested assessment rubric for depth studies is also provided.

Equilibrium and acid reactions

Outcomes

By the end of this module you will be able to:

- select and process appropriate qualitative and quantitative data and information using a range of appropriate media CH12-4
- analyse and evaluate primary and secondary data and information CH12-5
- solve scientific problems using primary and secondary data, critical thinking skills and scientific processes CH12-6
- communicate scientific understanding using suitable language and terminology for a specific audience or purpose CH12-7
- explain the characteristics of equilibrium systems, and the factors that affect these systems CH12-12

Content

STATIC AND DYNAMIC EQUILIBRIUM

INQUIRY QUESTION What happens when chemical reactions do not go through to completion?

By the end of this module you will be able to:

- conduct practical investigations to analyse the reversibility of chemical reactions, for example:
 - cobalt(II) chloride hydrated and dehydrated
 - iron(III) nitrate and potassium thiocyanate
 - burning magnesium
 - burning steel wool (ACSCH090) **ICT**
- model static and dynamic equilibrium and analyse the differences between open and closed systems (ACSCH079, ACSCH091)
- analyse examples of non-equilibrium systems in terms of the effect of entropy and enthalpy, for example:
 - combustion reactions
 - photosynthesis
- investigate the relationship between collision theory and reaction rate in order to analyse chemical equilibrium reactions (ACSCH070, ACSCH094) **ICT**

FACTORS THAT AFFECT EQUILIBRIUM

INQUIRY QUESTION What factors affect equilibrium and how?

By the end of this module you will be able to:

- investigate the effects of temperature, concentration, volume and/or pressure on a system at equilibrium and explain how Le Châtelier's principle can be used to predict such effects, for example:
 - heating cobalt(II) chloride hydrate
 - interaction between nitrogen dioxide and dinitrogen tetroxide
 - iron(III) thiocyanate and varying concentration of ions (ACSCH095)

Module 5 • Equilibrium and acid reactions

- explain the overall observations about equilibrium in terms of the collision theory (ACSCH094)
- examine how activation energy and heat of reaction affect the position of equilibrium

CALCULATING THE EQUILIBRIUM CONSTANT (K_{eq})

INQUIRY QUESTION How can the position of equilibrium be described and what does the equilibrium constant represent?

By the end of this module you will be able to:

- deduce the equilibrium expression (in terms of K_{eq}) for homogeneous reactions occurring in solution (ACSCH079, ACSCH096) **ICT N**
- perform calculations to find the value of K_{eq} and concentrations of substances within an equilibrium system, and use these values to make predictions on the direction in which a reaction may proceed (ACSCH096) **ICT N**
- qualitatively analyse the effect of temperature on the value of K_{eq} (ACSCH093) **ICT N**
- conduct an investigation to determine K_{eq} of a chemical equilibrium system, for example:
 - K_{eq} of the iron(III) thiocyanate equilibrium (ACSCH096) **ICT**
- explore the use of K_{eq} for different types of chemical reactions, including but not limited to:
 - dissociation of ionic solutions (ACSCH098, ACSCH099)

SOLUTION EQUILIBRIA

INQUIRY QUESTION How does solubility relate to chemical equilibrium?

By the end of this module you will be able to:

- describe and analyse the processes involved in the dissolution of ionic compounds in water
- investigate the use of solubility equilibria by Aboriginal and Torres Strait Islander Peoples when removing toxicity from foods, for example: **AHC**
 - toxins in cycad fruit
- conduct an investigation to determine solubility rules, and predict and analyse the composition of substances when two ionic solutions are mixed, for example:
 - potassium chloride and silver nitrate
 - potassium iodide and lead nitrate
 - sodium sulfate and barium nitrate (ACSCH065)
- derive equilibrium expressions for saturated solutions in terms of K_{sp} and calculate the solubility of an ionic substance from its K_{sp} value **ICT N**
- predict the formation of a precipitate given the standard reference values for K_{sp}

Key knowledge

Static and dynamic equilibrium

CHEMICAL SYSTEMS

In chemistry, a **system** is usually regarded as the chemical reaction being studied, and the **surroundings** are everything else (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.

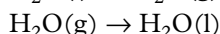
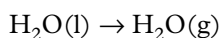
Systems are said to be **homogeneous** if all reactants and products are in the same phase, or **heterogeneous** if different phases are present.

Reversible reactions

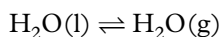
Chemical reactions that can occur in both directions are **reversible**. Chemical reactions that can occur only in one direction are **irreversible** or non-reversible, e.g. baking a cake.

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. Four examples of systems that can be reversible are:

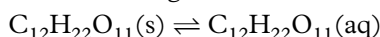
- the evaporation and condensation of water:



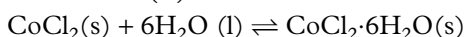
The reactions above can be written as:



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

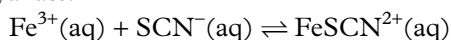


- the reaction between hydrated cobalt(II) chloride and dehydrated cobalt(II) chloride:



This reaction involves a colour change: CoCl_2 is blue and $\text{CoCl}_2 \cdot 6\text{H}_2\text{O}$ is pink.

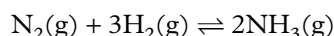
- the reaction of iron(III) nitrate and potassium thiocyanate:



This reaction also involves a colour change:

Fe^{3+} solutions are pale yellow, SCN^{-} solutions are colourless, and FeSCN^{2+} solutions are red.

Consider the reaction that occurs when N_2 gas is mixed with H_2 gas in a sealed container. NH_3 gas is formed, which then decomposes back to N_2 and H_2 . The reaction can be represented using double-headed arrows as:



When the rates of the forward and reverse reactions are equal the reaction is described as being in **equilibrium** (Figure 5.1).

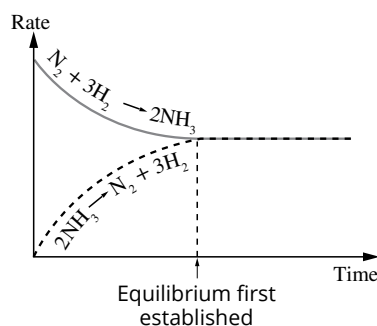


FIGURE 5.1 Graph showing the changes in the rates of the forward and reverse reactions with time as a reaction reaches equilibrium

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 by the concentration–time graph in Figure 5.2.

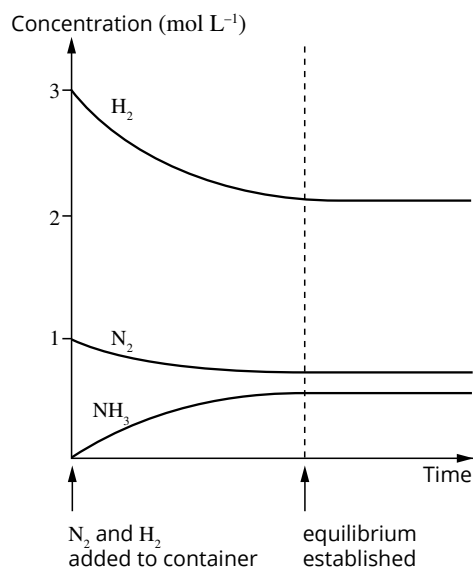


FIGURE 5.2 Changes in concentrations of N_2 , H_2 and NH_3 as a mixture of nitrogen and hydrogen gases react at a constant temperature and reach equilibrium

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

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. An energy profile diagram shows that once the products form, the reverse process can occur, in which product particles absorb energy to break their bonds and re-form the reactants. **GO TO** Year 11 Module 4

In Figure 5.3 the forward reaction is exothermic and the reverse reaction is endothermic. The activation energy for the endothermic reaction is the sum of the ΔH for the reaction and the activation energy for the exothermic reaction.

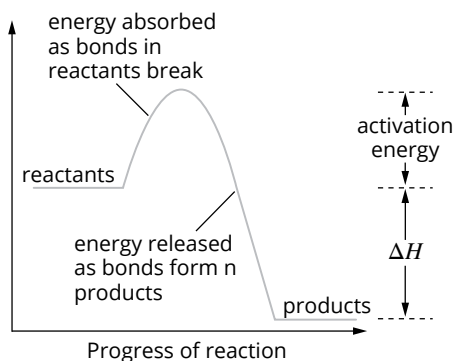


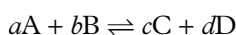
FIGURE 5.3 Energy profile diagram for an exothermic reaction

DYNAMIC EQUILIBRIUM

At equilibrium, the rate of the forward reaction is equal to the rate of the reverse reaction. But 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.

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:



is given by:

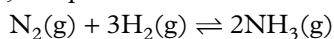
$$K_{\text{eq}} = \frac{[C]^c [D]^d}{[A]^a [B]^b}$$

The equilibrium constant, K_{eq} , 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 the concentrations of products to reactants can be calculated at any time during the reaction and is called the **reaction quotient**, Q , or concentration fraction.

- If $Q > K_{\text{eq}}$ the reaction shifts left to establish equilibrium.
- If $Q < K_{\text{eq}}$ the reaction shifts right to establish equilibrium.
- If $Q = K_{\text{eq}}$ the reaction is at equilibrium.

For example, at equilibrium for the reaction:



$$Q = K_{\text{eq}} = \frac{[\text{NH}_3]^2}{[\text{N}_2][\text{H}_2]^3}$$

You can understand the formation of the equilibrium using **collision theory**, which states that a reaction can occur when reactant particles collide with sufficient energy and have an appropriate molecular orientation.

When the reactants N_2 and H_2 are mixed, initially their particles collide and the reaction commences, producing particles of the product, NH_3 . Then, as the concentration of product particles increases and the concentration of reactant particles decreases, the rate of formation of product will decrease and the rate of reformation of reactants will increase. Eventually the rates of the forward and reverse reaction become equal.

The extent of reaction

The value of K_{eq} gives an indication of the **extent of the chemical reaction**, or the position of equilibrium. Different reactions proceed to different extents:

- If $K_{\text{eq}} < 10^{-4}$ there is a negligible forward reaction, and mainly reactants are present at equilibrium. The position of equilibrium is 'to the left'.
- If $10^{-4} < K_{\text{eq}} < 10^4$ the equilibrium mixture consists of significant amounts of both reactants and products.
- If $K_{\text{eq}} > 10^4$ there is an extensive forward reaction, and mainly products are present at equilibrium. The position of equilibrium is 'to the right'.

Factors affecting the value of K_{eq}

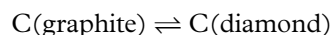
The value of the equilibrium constant for a reaction depends on:

- the coefficients in the equation: doubling the coefficients will square the original value of K_{eq} , while halving the coefficients will cause K_{eq} to be the square root of the original
- the direction of the equation: reversing the equation causes K_{eq} to have the inverse value
- the temperature: the effect of change of temperature on the value of K_{eq} is dependent on whether the reaction is exothermic or endothermic (discussed later, see Table 5.3).

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

STATIC EQUILIBRIUM

Static equilibrium in nature occurs when all forces are in balance; there is no overall force and there is no movement. In chemistry, the term 'static equilibrium' is used to describe chemical reactions in which there is no further conversion of reactants to products or products to reactants and the rates of the forward and reverse reactions are equal to zero. An example of static equilibrium could be considered to be:



Because the activation energy of the reaction for the conversion of diamond to graphite is very high, under normal conditions this reaction would take billions of years. As a result, the rates of the forward and reverse reactions are effectively zero. Table 5.1 summarises the differences between dynamic and static equilibrium in chemical systems.

TABLE 5.1 Differences between dynamic and static equilibrium in chemical systems

Static equilibrium	Dynamic equilibrium
Rates of forward and reverse reactions are equal and almost zero.	Rates of forward and reverse reactions are equal and not zero.
There is no movement of reactant and product particles in either the forward or reverse direction.	Although concentrations do not change, there is an equal rate of reaction of reactants to form products, and vice versa.

NON-EQUILIBRIUM SYSTEMS

Non-equilibrium chemical systems involve reactions that can be considered irreversible and never reach equilibrium. Two examples are combustion and photosynthesis.

For the combustion of hydrocarbons at standard conditions, ΔG° can be calculated using the equation:

$$\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ$$

ΔH° is negative and ΔS° is positive, so ΔG° is negative, indicating that combustion reactions are spontaneous systems. The products of the reaction, carbon dioxide and water, never recombine to form the hydrocarbon and oxygen so the reaction is described as irreversible.

For photosynthesis, ΔH° is positive, ΔS° is negative and ΔG° is positive. Photosynthesis is therefore a non-spontaneous, endothermic process. It is driven by, or coupled to, spontaneous reactions. The energy released by the spontaneous reactions is used during the many electron transfer steps that occur in photosynthesis.

Heat from these spontaneous reactions is also lost to the surroundings, causing an increase in entropy. Overall, the increase in the entropy of the universe caused by the spontaneous reactions is greater than the decrease in entropy caused by the production of glucose.

Factors that affect equilibrium

LE CHÂTELIER'S PRINCIPLE

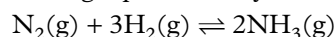
Le Châtelier's principle states: 'If a change is imposed on a system at equilibrium, the system will adjust itself to partially oppose the effect of the change'. The equilibrium will not completely return to its original state, but it will tend to oppose the effect.

The effect of a change on an equilibrium can be predicted by Le Châtelier's principle and understood in terms of collision theory and rates of reaction.

Concentration-time graphs and rate-time graphs

Concentration-time graphs and rate-time graphs help to understand the effects of changes to equilibrium.

Consider the following equilibrium system:



If extra nitrogen gas is added to the container without changing 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 5.4(a) is a concentration-time graph showing this effect.

The rate-time graph in Figure 5.4(b) shows the effects on the rate of the forward and reverse reactions when more nitrogen is added and the system returns to equilibrium. Initially there is a forward reaction. Then the rates of the forward and reverse reaction eventually become equal. The equilibrium position has shifted toward the right.

Because the temperature has not changed, the value of K_{eq} for the equilibrium reaction remains unchanged.

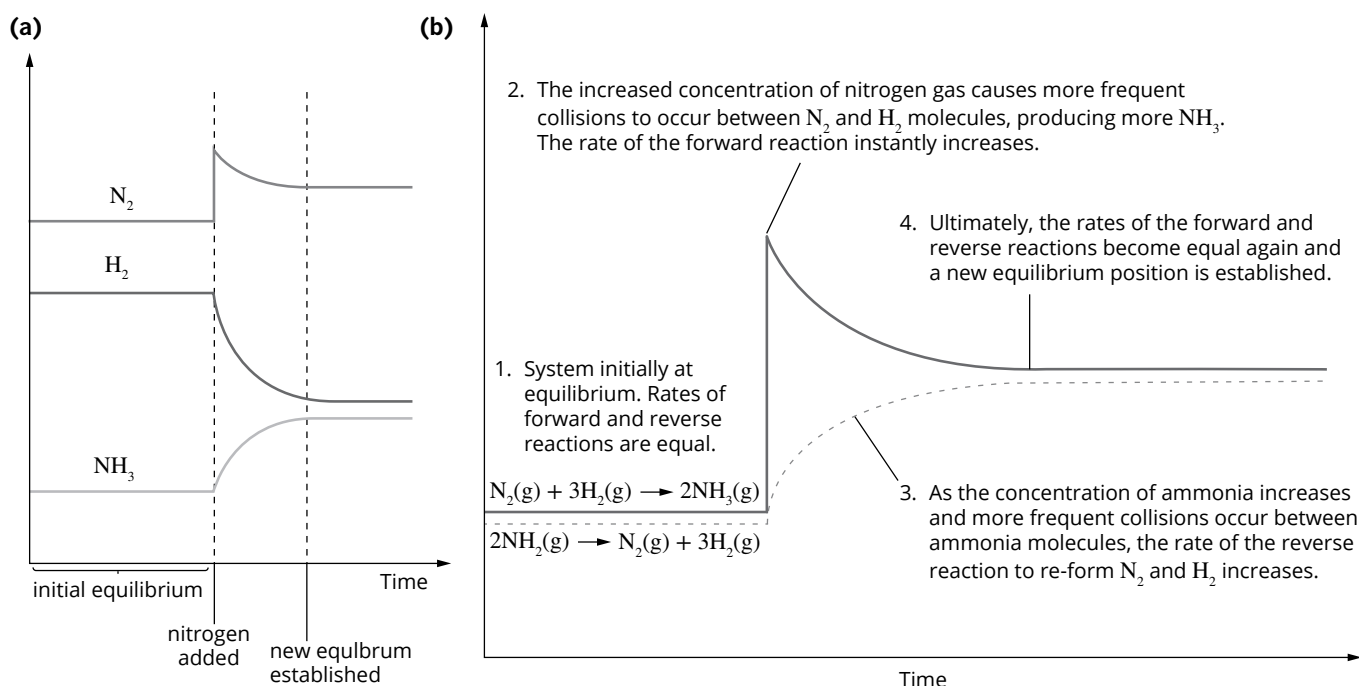


FIGURE 5.4 Concentration-time and rate-time graphs. (a) A representation of changes in concentrations that occur when additional nitrogen gas is added. (b) The effects on the rate of the forward and reverse reactions when more nitrogen is added.

Now consider the aqueous equilibrium system:

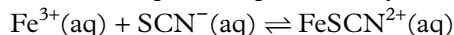


Figure 5.5 shows the effect of doubling the volume of the system by adding water without any change in temperature.

Although the equilibrium position shifts to the left, the concentrations of Fe^{3+} and SCN^{-} at the new equilibrium are lower than their concentrations prior to dilution, as the shift in the equilibrium only partially opposes the change.

Adding a catalyst speeds up the rate of the forward and backward reactions equally, so no change in the equilibrium concentrations of the reactants or products occurs. Hence, addition of a catalyst causes no change to the position of the equilibrium and no change in K_{eq} . Table 5.2 summarises the effect of various changes on an equilibrium system when the temperature is kept constant.

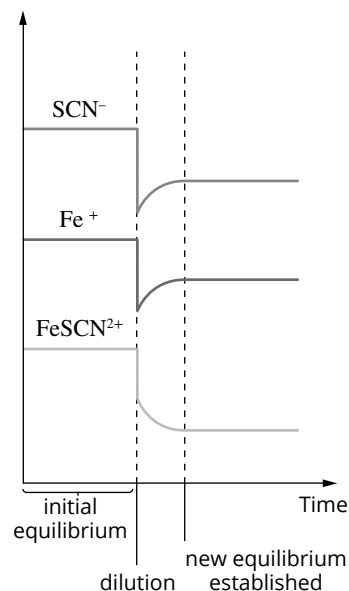


FIGURE 5.5 Effect of dilution on the equilibrium $\text{Fe}^{3+}(\text{aq}) + \text{SCN}^{-}(\text{aq}) \rightleftharpoons \text{FeSCN}^{2+}(\text{aq})$

TABLE 5.2 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 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 the 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) occurs to a greater extent.
Add a catalyst.	No change.	No change in either direction, because the rates of the forward and reverse reactions are increased equally. Frequencies of successful collisions are equal.
Add an inert gas (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) occurs to a greater extent.

Effect of temperature change on equilibrium

Only a change in temperature changes the value of K_{eq} . 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 5.3). These changes are described in Table 5.4 in terms of Le Châtelier's principle and collision theory.

Both the activation energy and heat of reaction affect the position of equilibrium. When the temperature increases, the rates of the forward and reverse reactions increase. The endothermic reaction of an equilibrium system 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 5.6). 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.

TABLE 5.3 The relationship between K_{eq} and ΔH as temperature increases

Sign of ΔH	Effect on K_{eq} as temperature increases
- (exothermic)	decreases
+ (endothermic)	increases

TABLE 5.4 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 effect of change on equilibrium position
Increasing the temperature for exothermic reactions	Shifts to the left (to absorb some of the added energy) and K_{eq} decreases.	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) and K_{eq} increases.	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.

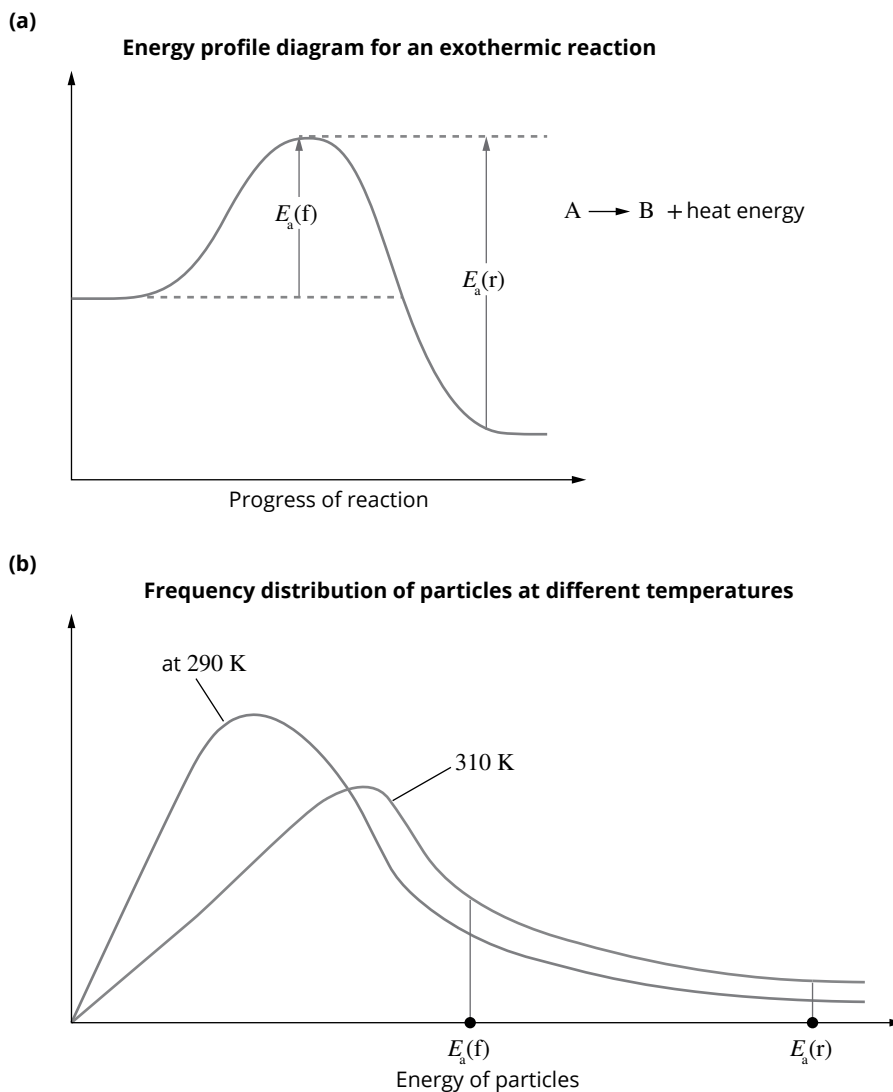
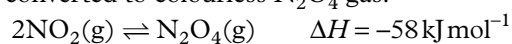


FIGURE 5.6 (a) Energy profile diagram for an exothermic reaction. The activation energy, $E_a(f)$, is less than the activation energy of the reverse reaction, $E_a(r)$. (b) The frequency distribution of molecules at two different temperatures. At the higher temperature, a greater proportion of particles 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:



An increase in temperature will cause the reverse reaction to occur, as indicated in Figure 5.7. The value of the K_{eq} for the new equilibrium will decrease.

In summary, as temperature increases, K_{eq} will:

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

The magnitude of ΔH determines the size of the change in K_{eq} for a particular change of temperature.

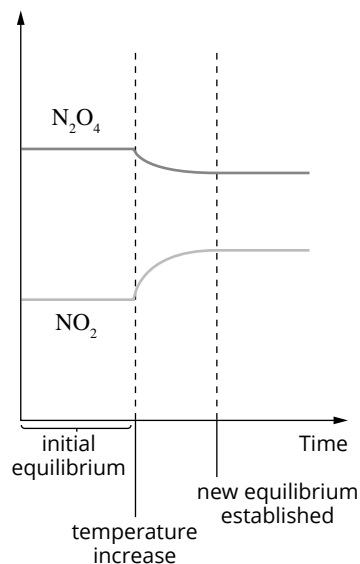


FIGURE 5.7 The effect of increasing the temperature on the equilibrium $2\text{NO}_2(\text{g}) \rightleftharpoons \text{N}_2\text{O}_4(\text{g})$

Calculating an equilibrium constant

At a specified temperature, the value of K_{eq} 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_{eq} and the concentrations of the other reactants and products are known.

The worksheets in this skills and assessment book will give you practice performing calculations involving equilibrium constants.

Solubility and equilibria

DISSOLUTION OF IONIC COMPOUNDS

The polar nature of water

Water is abundant in the environment. Its unique properties allow it to support life in many different ways. The key points about the structure of water are listed below.

- Each water molecule contains two hydrogen atoms covalently bonded to a single oxygen atom.
- The oxygen atom in each water molecule has two non-bonding electron pairs ('lone pairs').
- The covalent bonds within the molecule are polar; the oxygen atom has a higher electronegativity than the hydrogen atoms. The oxygen has a greater share of the shared electron pair and so carries a negative partial charge. Each hydrogen atom carries a positive partial charge.
- Water molecules are V-shaped and polar.
- The main intermolecular forces between water molecules are **hydrogen bonds** between the partial positive charge on a hydrogen atom on one molecule and a non-bonding pair of electrons on the oxygen atom of another (Figure 5.8).

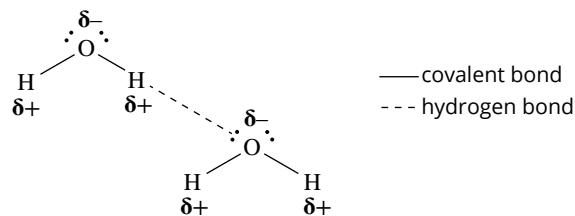


FIGURE 5.8 The structure of water molecules and the bonding between them

Water as a solvent

Solutes dissolve in solvents to form solutions. Figure 5.9 shows the three processes that occur when solutions form.

- In the solute, bonds between particles break.
- In the solvent, intermolecular bonds break.
- In the new solution, bonds form between solvent and solute particles.

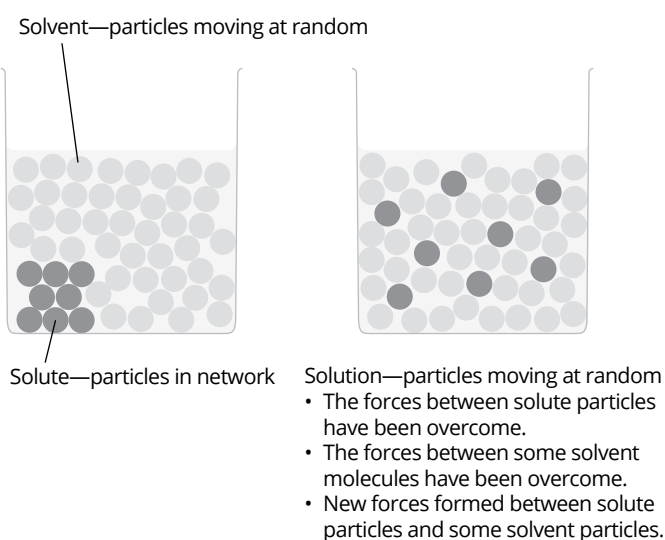


FIGURE 5.9 Rearrangement of particles when a solute dissolves in a solvent

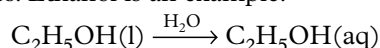
For a solute to dissolve, the force of attraction between the solute and solvent particles must be stronger than or similar to the solute-solute and solvent-solvent interparticle bonds.

The polar nature of water molecules and their ability to form hydrogen bonds enable water to dissolve other polar covalent substances as well as many ionic substances. Water does not readily dissolve non-polar substances.

When the maximum mass of solute dissolves at a specified temperature, the solution is described as **saturated**. **Unsaturated solutions** can dissolve more solute at the specified temperature. Supersaturated solutions are in an unstable state, in which more solute is dissolved than in a saturated solution at that temperature.

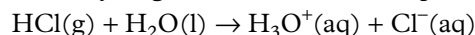
Water-soluble substances dissolve in one of three ways (Table 5.5):

- For polar covalent molecules that form hydrogen bonds, the polar molecules of the solute separate from each other and form hydrogen bonds with water molecules. Ethanol is an example.



- For polar molecules that ionise, the hydrogen atom has such a strong attraction to water that the polar covalent bond between the hydrogen atom and the rest of the molecule breaks (**dissociates**) and forms ions (**ionises**). The bonding electrons stay with the more electronegative atom, giving it a negative charge, and the hydrogen ion, or proton, bonds to water, forming a hydronium ion, H_3O^+ . The resultant ions then become **hydrated**, that is, surrounded by water

molecules. Hydrogen chloride is an example.



- For ionic compounds, attraction to the polar water molecules causes the ions in some ionic lattices to dissociate and become hydrated. Sodium chloride is an example:

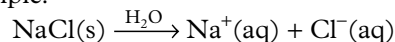


Table 5.6 summarises the different ways in which substances dissolve.

TABLE 5.5 Examples of the three ways solutes can dissolve in water

Substance	Bond broken in solute	Bonds formed with water	Arrangement of dissolved particles in solution
ethanol, $\text{C}_2\text{H}_5\text{OH}$	hydrogen bonds between ethanol molecules	hydrogen bonds between ethanol and water molecules	
hydrogen chloride, HCl	covalent bond between H and Cl atom in the hydrogen chloride molecule	covalent bond between hydrogen atom from HCl and oxygen atom on water molecule forming H_3O^+ ion ion-dipole bonds between newly formed ions and polar water molecules	<p>This attraction leads to the formation of a new covalent bond.</p> <p>This covalent bond breaks.</p>
sodium chloride, NaCl	ionic bonds between Na^+ and Cl^- ions	ion-dipole interactions between dissociated ions and polar water molecules	<p>----- Represents ion-dipole interaction</p>

TABLE 5.6 Summary of the different ways substances dissolve

	Ionic compounds	Non-ionising covalent compounds	Ionising covalent compounds
Examples	CuSO_4 , NaCl, Ca(OH)_2	sugar, glucose, ethanol	HCl, H_2SO_4 , HNO_3 , CH_3COOH
Name of process	dissociation	—	dissociation and ionisation
Particles present before dissolving occurs	ions	molecules	molecules
Particles present after dissolving occurs	ions	molecules	ions
Example of equation	$\text{NaCl(s)} \xrightarrow{\text{H}_2\text{O}} \text{Na}^+(\text{aq}) + \text{Cl}^-(\text{aq})$	$\text{C}_{12}\text{H}_{22}\text{O}_{11}\text{(s)} \xrightarrow{\text{H}_2\text{O}} \text{C}_{12}\text{H}_{22}\text{O}_{11}\text{(aq)}$	$\text{HCl(g)} + \text{H}_2\text{O(l)} \rightarrow \text{H}_3\text{O}^+(\text{aq}) + \text{Cl}^-(\text{aq})$

An application of solutions: Detoxifying food

For thousands of years, Aboriginal and Torres Strait Islanders have used different methods to remove toxicity from the cycad seeds, enabling them to be used as a rich food source. One simple way involved the solvent properties of water. The kernels were cut open and left in water so that the soluble toxins dissolved. Once the toxins were removed, the kernels were ground into flour and used to make bread.

SOLUBILITY OF IONIC SUBSTANCES

Not all ionic substances dissolve well in water, although their **solubility** tends to increase with increasing water

temperature. Solubility can be measured using different units, but mol L^{-1} is common. The solubility table below (Table 5.7) summarises the solubility of many common ionic compounds in water. It is useful to remember that all compounds containing the following ions are soluble:

- sodium
- nitrate
- ammonium
- acetate (also known as ethanoate)
- potassium.

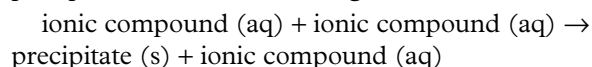
You can remember this as the SNAAP rule (**S**odium, **N**itrate, **A**mmonium, **A**cetate, **P**otassium). (You may also see this written as the SNAPE rule: Sodium, Nitrate, Ammonium, Potassium, Ethanoate.)

TABLE 5.7 Solubility of selected ionic compounds in water

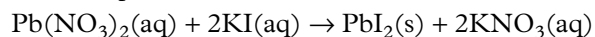
Soluble in water (> 0.1 mol dissolves per L at 25°C)	Exceptions: insoluble (< 0.01 mol dissolves per L at 25°C)	Exceptions: slightly soluble (0.01–0.1 mol dissolves per L at 25°C)
all nitrates, NO_3^-	no exceptions	no exceptions
all salts of ammonium, NH_4^+	no exceptions	no exceptions
all salts of sodium, Na^+ , and potassium, K^+	no exceptions	no exceptions
all ethanoates (acetates) CH_3COO^-	no exceptions	no exceptions
most sulfates, SO_4^{2-}	BaSO_4 , PbSO_4 , SrSO_4	CaSO_4 , Ag_2SO_4
most chlorides, Cl^- , bromides, Br^- , and iodides, I^-	AgCl , AgBr , AgI , PbI_2	no exceptions
Insoluble in water	Exceptions: soluble	Exceptions: slightly soluble
most hydroxides, OH^-	NaOH , KOH , $\text{Ba}(\text{OH})_2$, NH_4OH	$\text{Ca}(\text{OH})_2$, $\text{Sr}(\text{OH})_2$
most carbonates, CO_3^{2-}	Na_2CO_3 , K_2CO_3 , $(\text{NH}_4)_2\text{CO}_3$	No important exceptions
most phosphates, PO_4^{3-}	Na_2PO_4 , K_2PO_4 , $(\text{NH}_4)_2\text{PO}_4$	No important exceptions
most sulfides, S^{2-}	Na_2S , K_2S , $(\text{NH}_4)_2\text{S}$	No important exceptions

Precipitation reactions

When two aqueous solutions are mixed, an insoluble substance called a **precipitate** sometimes forms (Figure 5.10). The balanced chemical equation for a precipitation reaction has the general form:



For example:



In this case, PbI_2 is the precipitate. K^+ and NO_3^- ions remain dissolved in solution and are termed **spectator ions**.

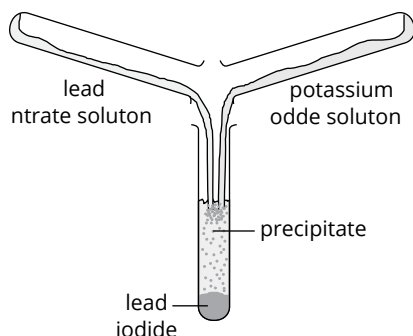
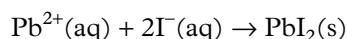


FIGURE 5.10 Mixing solutions of lead nitrate and potassium iodide forms a precipitate of solid lead iodide.

An ionic equation omits spectator ions to give a more accurate picture of the reaction taking place. For example:



When writing balanced ionic equations, ensure that:

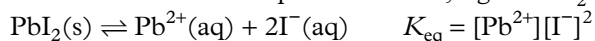
- species that have (s), (l) or (g) beside them appear in the equation and are not separated into different particles.
- ions with (aq) beside them that appear on both sides of the arrow are omitted from the equation. These are spectator ions.

SOLUBILITY AND EQUILIBRIUM

K_{eq} for heterogeneous equilibria

The expression for K_{eq} for a homogeneous equilibrium includes the molar concentrations (concentrations in mol L^{-1}) for all the reactants and products in the equilibrium. However, 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_{eq} . These concentrations are constant and can be removed from the expression for the K_{eq} .

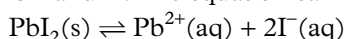
An example of a heterogenous equilibrium and the associated expression for K_{eq} is the dissolution of an ionic solid in water to form an aqueous solution, e.g for PbI_2 :



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 5.11 where equilibrium exists between the solid and its ions, Pb^{2+} and I^- . The equation can be written as:



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

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

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.

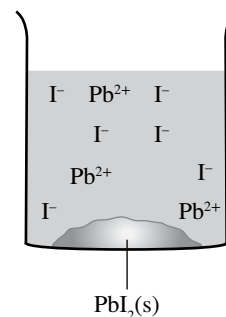


FIGURE 5.11 A saturated solution of solid PbI_2 in equilibrium with its ions, Pb^{2+} and I^-

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 5.8 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 (Table 5.9) by calculating the reaction quotient, Q , for the system. For the lead iodide solution, Q is equal to $[Pb^{2+}][I^-]^2$. Q is often called the **ionic product** when referring to the dissolution of ionic solids in solution.

TABLE 5.8 Conversion from solubility, s , in $mol\ L^{-1}$ to K_{sp} where s 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 5.9 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 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 solid is present, some solid dissolves to establish equilibrium.

WORKSHEET 5.1

Knowledge review—thinking about rates and energy

- 1 Complete the following table by providing the term to match the definition. 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 part of the universe we are studying.	
The name for a theory that accounts for the rates of chemical reactions.	
A system in which only energy is exchanged with the surroundings.	
A substance that increases the rate of reaction but is not consumed in the reaction.	
The speed at which the reaction occurs.	
A reaction that releases energy to the surroundings.	
Name for a chemical equation that includes the enthalpy change of the reaction.	
A reaction for which the ΔH is positive.	
A measure of the number of possible arrangements of a system or the degree of disorder or randomness in the system.	
The quantity that is used to determine if a reaction is spontaneous.	

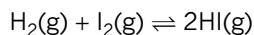
- 2 Decide whether each of the following statements is true or false.

Statement	True/False
Rates of reactions can be increased by increasing the concentration of solutions, increasing the surface area of solid reactants and increasing the temperature of the system.	
When ΔG is greater than zero the reaction is non-spontaneous.	
Reactions that are exothermic (ΔH negative) and have an increase in entropy (ΔS positive) are always spontaneous.	
Open systems are only able to exchange matter with the surroundings.	

WORKSHEET 5.2

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 gives the results of seven of these experiments for the following equilibrium:

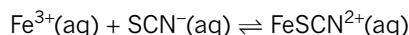


Equilibrium concentrations (mol L ⁻¹) for $\text{H}_2(\text{g}) + \text{I}_2(\text{g}) \rightleftharpoons 2\text{HI}(\text{g})$ at 458°C						
Mixture	[H ₂]	[I ₂]	[HI]	[H ₂] + [I ₂] + [HI]	$\frac{[\text{HI}]}{[\text{H}_2][\text{I}_2]}$	$\frac{[\text{H}]^2}{[\text{H}_2][\text{I}_2]}$
1	0.0056170	0.0005940	0.0126990			
2	0.0045800	0.0009930	0.0148580			
3	0.0038420	0.0015240	0.0168710			
4	0.0046670	0.0010580	0.0154450			
5	0.0016960	0.0016960	0.0118070			
6	0.0014330	0.0014330	0.0099990			
7	0.0042130	0.0042130	0.0294350			

- a Calculate the missing entries in the blank columns in the spreadsheet. Which expression was almost constant?
- _____
- b The ratio that was almost constant is called the equilibrium constant, K_{eq} , for this reaction at the specified temperature. Write the mathematical expression for K_{eq} for this reaction.
- _____
- c Write a general expression for the equilibrium law for the following reaction:
- $$a\text{A}(\text{g}) + b\text{B}(\text{g}) \rightleftharpoons c\text{C}(\text{g}) + d\text{D}(\text{g})$$
- _____
- 2 a Write the equation $\text{H}_2(\text{g}) + \text{I}_2(\text{g}) \rightleftharpoons 2\text{HI}(\text{g})$ in reverse. _____
- b Write the mathematical 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_{eq} for this reverse reaction.
- _____
- d What is the relationship between the constant calculated in part c and the constant calculated in question 1?
- _____
- e Use your answer to part c to calculate the value of K_{eq} for the following reaction:
- $$\frac{1}{2}\text{H}_2(\text{g}) + \frac{1}{2}\text{I}_2(\text{g}) \rightleftharpoons \text{HI}(\text{g})$$
- _____
- f What is the relationship between the constant you calculated in part e and the constant you calculated in question 1?
- _____
- g Complete the following summary statements:
- If $K_{\text{eq}1}$ is the equilibrium constant for an equation, the constant for the reverse equation is given by $K_{\text{eq}2} = \text{_____}$.
- If the coefficients in the equation are halved, the constant for the new equation is given by $K_{\text{eq}3} = \text{_____}$.

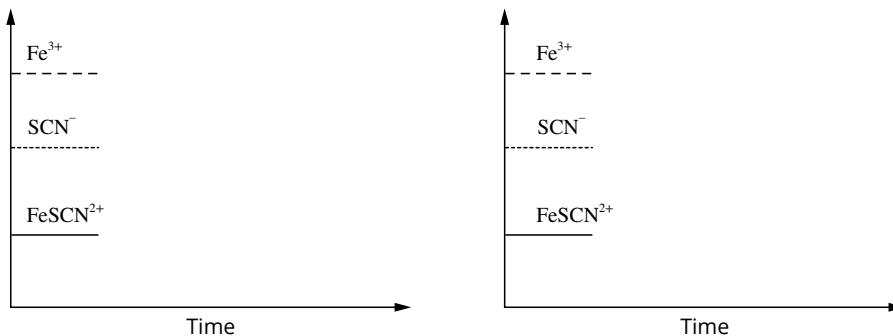
WORKSHEET 5.2

3 a Consider the following solution equilibrium:

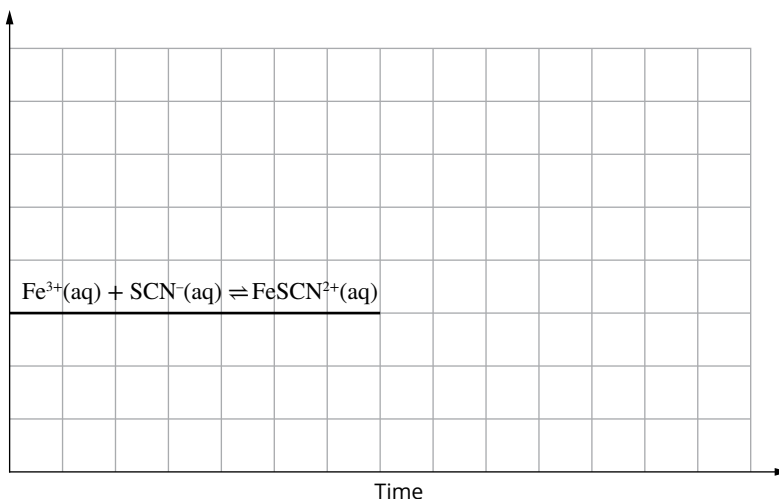


Sketch concentration–time graphs on the grids provided below if each of the following changes were made to this system at constant temperature.

- i A small amount of $\text{Fe}^{3+}(\text{aq})$ is added. (Assume the volume change is negligible.)
- ii A volume of water is added to the system.



b Sketch the rate–time graph for this system when a small amount of $\text{SCN}^{-}(\text{aq})$ is added, assuming no volume change. Your graph should show what happens to the rates of both the forward and reverse reactions.



c In terms of collision theory, explain the effects on the rates of the forward and reverse reactions when more SCN^{-} is added and as the system returns to equilibrium.

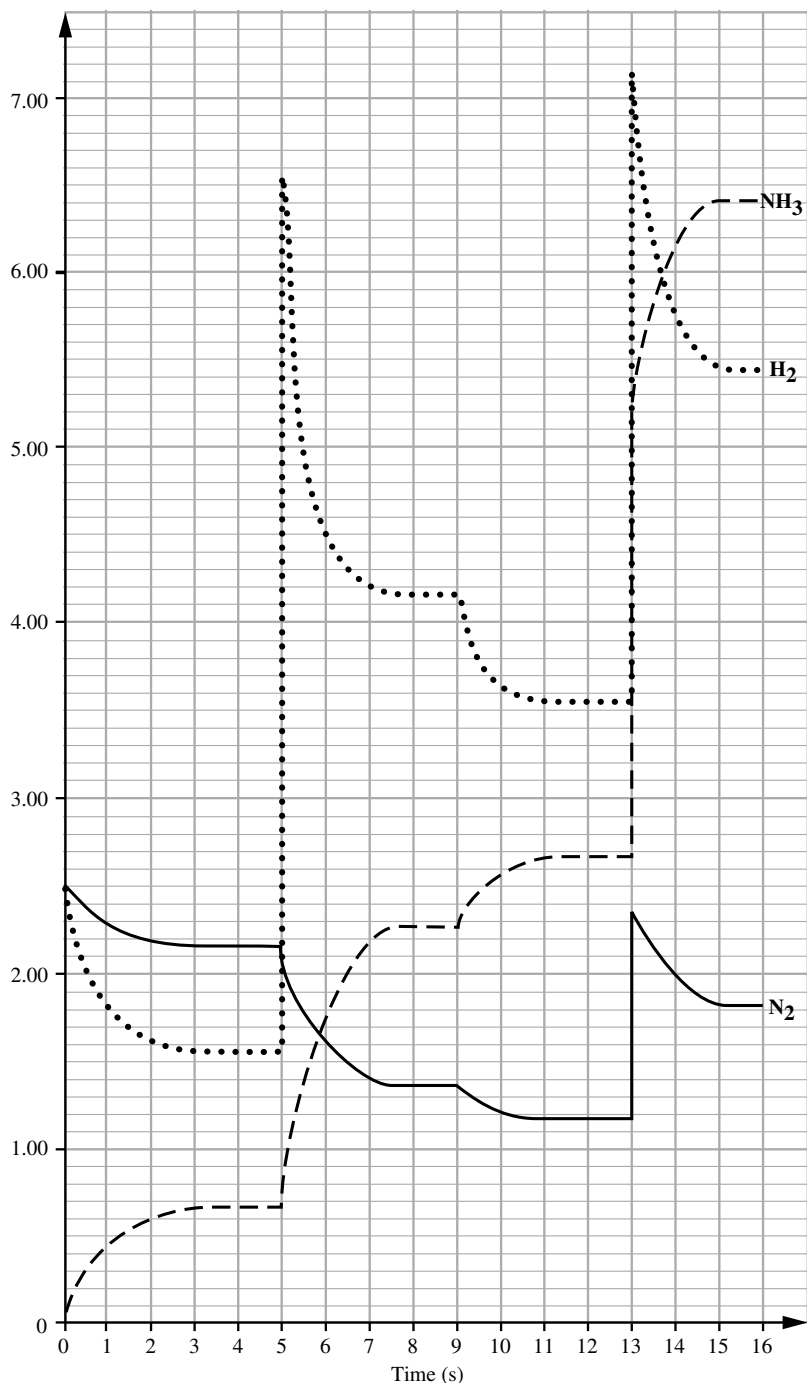
4 The following table shows the results of a simple experiment involving temperature changes on the equilibrium at constant volume.

Equilibrium system	Type of reaction	Colour in ice bath	Colour in hot water bath
$\text{Fe}^{3+}(\text{aq}) + \text{SCN}^{-}(\text{aq}) \rightleftharpoons \text{FeSCN}^{2+}(\text{aq})$ (pale yellow) (red)	exothermic	red	pale yellow

Complete the table below.

Equilibrium system	Effect of heating on concentration of products	After heating, the position of equilibrium has moved ...	Effect of heating on value of K_{eq}
$\text{Fe}^{3+}(\text{aq}) + \text{SCN}^{-}(\text{aq}) \rightleftharpoons \text{FeSCN}^{2+}(\text{aq})$ (pale yellow) (red)			

5 Examine the graph for the following reaction, and answer the questions for the system: $\text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightleftharpoons 2\text{NH}_3(\text{g})$



a During which time intervals was the system at equilibrium? _____

b Calculate the equilibrium constant at: (i) 4 seconds, (ii) 8 seconds, (iii) 12 seconds, and (iv) 16 seconds.

c Using the answers in part b and your knowledge of equilibrium, explain what changes were made to the system at: (i) 5 seconds, (ii) 9 seconds, and (iii) 13 seconds.

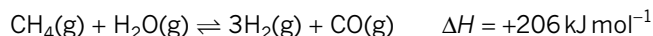
RATING MY LEARNING	My understanding improved	Not confident	←	→	Very confident	I answered questions without help	Not confident	←	→	Very confident	I corrected my errors without help	Not confident	←	→	Very confident
	<input type="radio"/>	<input type="radio"/>	<input type="radio"/>	<input type="radio"/>	<input type="radio"/>		<input type="radio"/>	<input type="radio"/>	<input type="radio"/>	<input type="radio"/>		<input type="radio"/>	<input type="radio"/>	<input type="radio"/>	<input type="radio"/>

WORKSHEET 5.3

Calculations—equilibrium constants and concentrations

This worksheet allows you to practise doing calculations involving chemical equilibria.

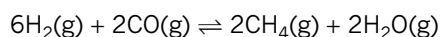
Steam reforming is used for the large-scale industrial production of hydrogen gas. In this process, methane gas is converted to synthesis gas, which is a mixture of carbon monoxide gas and hydrogen gas. The thermochemical equation is:



The equilibrium constant at 650°C is 2.4×10^{-4} .

1 Write the expression for the equilibrium constant for the equation above.

2 Calculate the equilibrium constant at 650°C and the ΔH value for the equation below.



3 The following gases are added to a sealed 2.00 L container at a constant temperature of 650°C.

- 0.012 mol of methane gas
- 0.0080 mol of water vapour
- 0.016 mol of carbon monoxide gas
- 0.0060 mol of hydrogen gas

a Determine the initial value of the reaction quotient (concentration fraction), Q .

b In which direction would the reaction move to establish equilibrium? Explain.

c The temperature of the equilibrium is increased to 850°C.

i Will the equilibrium constant increase, decrease or remain unchanged? Explain your answer.

.....
WORKSHEET 5.3

ii How will the increased temperature affect the rate and equilibrium yield of the reaction? Explain your answer.

4 An equilibrium mixture of CH₄, H₂O, H₂ and CO was prepared in a 2.00L flask at another temperature by adding 0.600 mol CH₄ and 0.400 mol of H₂O to the reaction vessel. At equilibrium there was 0.110 mol of CO present.

a Complete the following ICE (Initial, Change, Equilibrium) table to determine the amounts of CH₄, H₂O and H₂ at equilibrium. Let *x* be the amount of CH₄ used.

	CH ₄	+ H ₂ O	⇌ 3H ₂	+ CO
Initial (mol)	0.600	0.400	0.000	0.000
Change (mol)				
Equilibrium (mol)				0.110

b Calculate the equilibrium concentrations of CH₄, H₂O, H₂ and CO, in mol L⁻¹.

c Calculate the equilibrium constant, *K*_{eq}, for the reaction at this temperature.

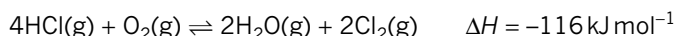
RATING MY LEARNING	My understanding improved	Not confident	←	→	Very confident	I answered questions without help	Not confident	←	→	Very confident	I corrected my errors without help	Not confident	←	→	Very confident
		<input type="radio"/>	<input type="radio"/>	<input type="radio"/>	<input type="radio"/>		<input type="radio"/>	<input type="radio"/>	<input type="radio"/>	<input type="radio"/>		<input type="radio"/>	<input type="radio"/>	<input type="radio"/>	<input type="radio"/>

WORKSHEET 5.4

Equilibrium—Le Châtelier's principle and the equilibrium law

- 1 State Le Châtelier's principle.

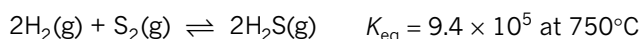
- 2 Consider the following equilibrium system:



Predict the effect (increase, decrease, no change) of each of the following changes to the equilibrium system on each quantity in the following table. Briefly explain the reason for your answer by using Le Châtelier's principle.

Predicted effects on equilibrium system			
Quantity	Change	Effect (compared to initial equilibrium)	Reason
[Cl ₂]	Add O ₂ at constant volume and temperature.		
amount of H ₂ O (in mol)	Decrease the volume of the container at constant temperature.		
[HCl]	Decrease the pressure by increasing the volume of the container at constant temperature.		
K _{eq}	Increase the volume of the container at constant temperature.		
K _{eq}	Increase the temperature at constant volume.		
[O ₂]	Increase the temperature at constant volume.		
[HCl]	Add argon gas at constant volume and temperature.		
K _{eq}	Add a catalyst at constant volume and temperature.		

- 3 Consider the following reaction:



A mixture of H₂, S₂ and H₂S was allowed to come to equilibrium in a closed 2.0L container at 750°C. The equilibrium concentrations of H₂ and H₂S gases were analysed and found to be 0.234 mol L⁻¹ and 0.442 mol L⁻¹ respectively.

- a What does the value of the equilibrium constant for this reaction tell you about the extent of the reaction?

- b Write an expression for the equilibrium constant for this reaction.

- c Calculate the equilibrium concentration of S₂(g) in the mixture.

.....
WORKSHEET 5.4

4 1.364 mol of H₂, 0.682 mol of S₂ and 0.680 mol of H₂S were mixed in another 2.00L container at 580°C. At equilibrium the concentration of H₂S was measured as 1.010 mol L⁻¹.

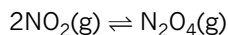
a What is the value of the equilibrium constant at this temperature?

b Compare your answer to the value of K_{eq} for this reaction in question 3. Is this reaction exothermic or endothermic? Give a reason for your answer.

5 Consider the following reaction at equilibrium:



a Calculate the equilibrium constant at 327°C for the following reaction:



b In terms of Le Châtelier's principle, explain the effect on the position of equilibrium when:

i the temperature is decreased at constant volume

ii the volume of the container is increased at constant temperature.

6 The value of ΔH for a chemical reaction depends on the direction of the equation for the reaction and the coefficients and states of species in the equation. These factors also affect the value of an equilibrium constant. However, when an equation is written in a different way the effects on ΔH and K_{eq} are different.

Complete the following table for the two reactions shown.

Effects on ΔH and K _{eq}					
Equation	ΔH (kJ mol ⁻¹)	K _{eq} at 527°C	Equation	ΔH (kJ mol ⁻¹)	K _{eq} at 600°C
N ₂ (g) + 3H ₂ (g) ⇌ 2NH ₃ (g)	-92	0.051	2SO ₂ (g) + O ₂ (g) ⇌ 2SO ₃ (g)	-198	1.56
2NH ₃ (g) ⇌ N ₂ (g) + 3H ₂ (g)				-99	
2N ₂ (g) + 6H ₂ (g) ⇌ 4NH ₃ (g)					2.43
NH ₃ (g) ⇌ ½N ₂ (g) + ¾H ₂ (g)				+198	

7 Use your results from question 6 to complete the summary statements below.

For ΔH: When equations are reversed, ΔH has the _____ sign and _____ magnitude.

When coefficients are doubled, ΔH has the _____ sign and the magnitude is _____.

When coefficients are halved, ΔH has the _____ sign and the magnitude is _____.

For K_{eq}: When equations are reversed, the value of K_{eq} is _____.

When coefficients are doubled, the value of K_{eq} is _____.

When the coefficients are halved, the value of K_{eq} is _____.

RATING MY LEARNING	My understanding improved	Not confident	←	→	Very confident	I answered questions without help	Not confident	←	→	Very confident	I corrected my errors without help	Not confident	←	→	Very confident
	<input type="radio"/>	<input type="radio"/>	<input type="radio"/>	<input type="radio"/>	<input type="radio"/>	<input type="radio"/>	<input type="radio"/>	<input type="radio"/>	<input type="radio"/>	<input type="radio"/>	<input type="radio"/>	<input type="radio"/>	<input type="radio"/>	<input type="radio"/>	<input type="radio"/>

WORKSHEET 5.5

Wonderful water—structure and properties

1 Draw the valence structures of five water molecules as they might appear in a sample of liquid water. Don't forget to show the non-bonding electron pairs (lone pairs) and intermolecular forces.

2 a What type of bonding holds the oxygen and hydrogen atoms together within the molecule? _____.
Label one of these bonds on your diagram in question 1.

b Explain why water is a polar molecule.

c The attraction between different water molecules is mainly due to hydrogen bonding. Label a hydrogen bond on your diagram in question 1.

d Explain what causes hydrogen bonding and why it is stronger than other types of dipole-dipole bonding.

3 a Write a balanced equation to show KI dissolving in water.

b Write a balanced equation for CH₃OH dissolving in water.

c Write a balanced equation for HCl gas dissolving in water.

4 Sketch the arrangement of water molecules around potassium and iodide ions when KI is dissolved in water.

RATING MY LEARNING	My understanding improved	Not confident	←	→	Very confident	I answered questions without help	Not confident	←	→	Very confident	I corrected my errors without help	Not confident	←	→	Very confident
	<input type="radio"/>	<input type="radio"/>	<input type="radio"/>	<input type="radio"/>	<input type="radio"/>		<input type="radio"/>	<input type="radio"/>	<input type="radio"/>	<input type="radio"/>		<input type="radio"/>	<input type="radio"/>	<input type="radio"/>	<input type="radio"/>

WORKSHEET 5.6

Solving solubility—predicting precipitation reactions and writing ionic equations

A solubility table summarises the solubility in water of common ionic compounds. Use the solubility data in Table 5.7 on page 10 to answer the questions in this worksheet.

- 1 The following solutions are mixed together. Indicate whether or not you think a precipitate will form by placing a tick (for yes) or a cross (for no) in each box in the second column. For each predicted precipitate, write a fully balanced chemical equation in the third column.

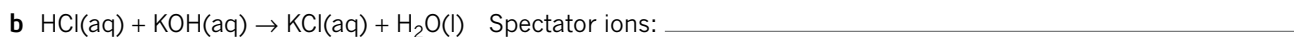
Solutions mixed	Precipitate? (✓ or ×)	Balanced chemical equation
potassium chloride + silver nitrate		
copper(II) nitrate + sodium hydroxide		
magnesium nitrate + sodium chloride		
lead nitrate + potassium iodide		
sodium sulfate + barium nitrate		

- 2 Outline an experimental method a student could follow to obtain a dry sample of silver chloride from solutions of potassium chloride and silver nitrate.

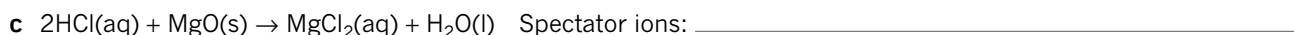
- 3 Identify the spectator ions in each of the reactions below, and write the ionic equation.



Ionic equation: _____



Ionic equation: _____



Ionic equation: _____

- 4 Use the internet to investigate and compare the processes used by Aboriginal and Torres Strait Islanders to remove toxins from foods. Present your findings in a short report below.

RATING MY LEARNING	My understanding improved	Not confident	←	→	Very confident	I answered questions without help	Not confident	←	→	Very confident	I corrected my errors without help	Not confident	←	→	Very confident
	<input type="radio"/>	<input type="radio"/>	<input type="radio"/>	<input type="radio"/>	<input type="radio"/>		<input type="radio"/>	<input type="radio"/>	<input type="radio"/>	<input type="radio"/>		<input type="radio"/>	<input type="radio"/>	<input type="radio"/>	<input type="radio"/>

WORKSHEET 5.7

Solubility equilibrium—calculating solubility and K_{sp}

A saturated solution forms when the rate of dissolution of an ionic solid equals the rate of precipitation of its ions. This dynamic equilibrium is independent of the amount of solid present, although there must be some solid in equilibrium with the solution. The equilibrium constant for this type of equilibrium is called the solubility product, K_{sp} .

- 1 For each of the following ionic solids, complete the table by writing the equation for its dissolution in water and the expression for the solubility products.

Solid	Equation for dissolution in water	Expression for K_{sp}
BaSO ₄		
AgCl		
Ag ₂ CO ₃		
Al(OH) ₃		

- 2 a The values of K_{sp} at 25°C for the solids in question 1 are listed in the table below. Calculate the solubility of each substance, in mol L⁻¹.

Solid	K_{sp}	Calculation of solubility, s (in mol L ⁻¹)
BaSO ₄	1.08×10^{-10}	
AgCl	1.77×10^{-10}	
Ag ₂ CO ₃	8.46×10^{-12}	
Al(OH) ₃	3.0×10^{-34}	

- b List these solids in order of solubility, from most soluble to least soluble.
- _____
- 3 Your teeth are composed mainly of the ionic compounds calcium phosphate, Ca₃(PO₄)₂ and hydroxyapatite, Ca₅(PO₄)₃(OH).
- a Given the solubility of Ca₃(PO₄)₂ is 7.19×10^{-7} , calculate the K_{sp} of Ca₃(PO₄)₂.
- b Explain why the consumption of carbonated drinks, which have quite a low pH, could increase dental decay.
- _____
- _____

RATING MY
LEARNING

My
understanding
improved

Not
confident

← →

Very
confident

I answered
questions
without help

Not
confident

← →

Very
confident

I corrected
my errors
without help

Not
confident

← →

Very
confident

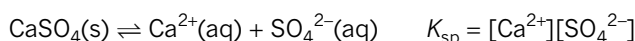
WORKSHEET 5.8

Predicting precipitation using K_{sp}

The reaction quotient, Q , for a system that reaches equilibrium can be determined at any stage during the reaction. By comparing the size of Q with the equilibrium constant, K_{eq} , you can determine which way the reaction will proceed to establish equilibrium, at which point $Q = K_{eq}$.

When an ionic solid forms a saturated solution, the solid is in equilibrium with its ions in solution and $Q = K_{sp}$, where Q is the ion product and K_{sp} is the solubility product. You can predict whether precipitation will occur when two solutions of ions are mixed by comparing Q and K_{sp} . In general, precipitation occurs if the ionic product is greater than the solubility product.

Remember that K_{sp} is the equilibrium constant for the reaction written as the solid in equilibrium with its ions, for example:



- 1 Consider the following mixtures of chemicals. Complete the table by writing balanced equations and calculating the value of Q . Hence indicate whether precipitation will occur.

Mixture	Balanced equation	K_{sp}	Q	Precipitation occurs (Y or N)?
100 mL 0.5 mol L ⁻¹ Ca(NO ₃) ₂ mixed with 100 mL 0.2 mol L ⁻¹ Na ₂ SO ₄		6.1×10^{-5}		
20 mL 0.01 mol L ⁻¹ AgNO ₃ mixed with 80 mL 0.05 mol L ⁻¹ K ₂ CrO ₄		1.7×10^{-12}		

- 2 For each of the following 1 L solutions, calculate the minimum concentration of lead ions, Pb²⁺, in the solution that is needed to cause precipitation of the stated compound.

Precipitate	K_{sp}	Solution	Balanced ionic equation	Calculation of [Pb ²⁺] (in mol L ⁻¹)
PbS	3×10^{-28}	0.01 mol L ⁻¹ Na ₂ S		
PbF ₂	3.3×10^{-8}	0.01 mol L ⁻¹ NaF		

- 3 A sample of bore water from a mine contains equal amounts, in mol, of dissolved silver carbonate and iron(II) carbonate. The values of the K_{sp} for these chemicals are:

$$K_{sp}(\text{Ag}_2\text{CO}_3) = 8.10 \times 10^{-12}$$

$$K_{sp}(\text{FeCO}_3) = 3.13 \times 10^{-11}$$

Using this data, determine the order in which the ionic solids will precipitate if the water is allowed to evaporate. (Show your calculations below.)

RATING MY LEARNING	My understanding improved	Not confident	←	→	Very confident	I answered questions without help	Not confident	←	→	Very confident	I corrected my errors without help	Not confident	←	→	Very confident
	<input type="radio"/>	<input type="radio"/>	<input type="radio"/>	<input type="radio"/>	<input type="radio"/>		<input type="radio"/>	<input type="radio"/>	<input type="radio"/>	<input type="radio"/>		<input type="radio"/>	<input type="radio"/>	<input type="radio"/>	<input type="radio"/>

WORKSHEET 5.9

Literacy review—equilibrium terms and expressions

1 Complete the following table to review your understanding of some of the terms in this module.

Term	Meaning
K_{eq}	
Q	
K_{sp}	
equilibrium law	
spectator ions	

2 Select the correct mathematical expressions below and arrange them in order to show how to calculate the solubility of $\text{Ag}_3(\text{PO}_4)_2$, given the value of the solubility product K_{sp} is 8.89×10^{-17} at 25°C .

a Let s be the solubility in mol L^{-1} of $\text{Ag}_3(\text{PO}_4)_2$

b $3s \times 2s = 8.89 \times 10^{-17}$

c $(3s)^3 \times (2s)^2 = 8.89 \times 10^{-17}$

d $[\text{Ag}^+]^3 \times [\text{PO}_4^{3-}]^2 = 8.89 \times 10^{-17}$

e $[\text{Ag}^+] = s$

f $[\text{Ag}^+] = 3s$

g $[\text{PO}_4^{3-}] = 2s$

h $[\text{Ag}^+] \times [\text{PO}_4^{3-}] = 8.89 \times 10^{-17}$

3 The following paragraphs about equilibrium contain several factual errors. Highlight the errors and correct them below.

Dynamic equilibrium occurs when there is a reversible reaction in an open system. In an equilibrium system the rates of the forward and reverse reactions are equal and there is no change in the concentrations of the reactant and product molecules. However, it can be shown that there is continual conversion of reactants to products and vice versa. If the system is closed, it will never achieve equilibrium because only energy is lost to the surroundings.

The equilibrium law states that at a specific temperature, the equilibrium constant, K_{eq} , is equal to only the concentration of the products divided by the concentration of the reactants. If this concentration ratio, Q , is less than K_{eq} the system will favour the reverse reaction to restore equilibrium.

When the temperature is decreased for an exothermic reaction, the value of K_{eq} decreases. When the temperature is increased for an endothermic reaction, K_{eq} increases.

When reactants are mixed in an open system, collision theory explains that initially, the rate of the forward reaction is low and decreases as the reactants' concentrations decrease. Simultaneously, the reverse reaction is very fast initially and increases as the concentrations of the products increase and collisions between product molecules increase.

Le Châtelier's principle predicts that when the pressure of a system decreases by increasing the volume, the system will move in the direction of the least particles in order to re-establish equilibrium.

RATING MY LEARNING	My understanding improved	Not confident	←	→	Very confident	I answered questions without help	Not confident	←	→	Very confident	I corrected my errors without help	Not confident	←	→	Very confident
	<input type="radio"/>	<input type="radio"/>	<input type="radio"/>	<input type="radio"/>	<input type="radio"/>	<input type="radio"/>	<input type="radio"/>	<input type="radio"/>	<input type="radio"/>	<input type="radio"/>	<input type="radio"/>	<input type="radio"/>	<input type="radio"/>	<input type="radio"/>	<input type="radio"/>

WORKSHEET 5.10

Thinking about my learning

On completion of Module 5: Equilibrium and acid reactions, you should be able to describe, explain and apply the relevant scientific ideas. You should also be able to interpret, analyse and evaluate data.

1 The table lists the key knowledge covered in this module. Read each and reflect on how well you understand each concept. Rate your learning by shading the circle that corresponds to your level of understanding for each concept. It may be helpful to use colour as a visual representation. For example:

- green—very confident
- orange—in the middle
- red—starting to develop.

Concept focus	Rate my learning				
	Starting to develop ←————→ Very confident				
Static and dynamic equilibrium	<input type="radio"/>	<input type="radio"/>	<input type="radio"/>	<input type="radio"/>	<input type="radio"/>
Factors that affect equilibrium	<input type="radio"/>	<input type="radio"/>	<input type="radio"/>	<input type="radio"/>	<input type="radio"/>
Calculating the equilibrium constant (K_{eq})	<input type="radio"/>	<input type="radio"/>	<input type="radio"/>	<input type="radio"/>	<input type="radio"/>
Solution equilibria	<input type="radio"/>	<input type="radio"/>	<input type="radio"/>	<input type="radio"/>	<input type="radio"/>

2 Consider points you have shaded from starting to develop to middle-level understanding. List specific ideas you can identify that were challenging.

3 Write down two different strategies that you will apply to help further your understanding of these ideas.

PRACTICAL ACTIVITY 5.1

Reversible and irreversible reactions

Suggested duration: 25 minutes

INTRODUCTION

An open system is a system that exchanges matter and energy with the surroundings. A closed system only exchanges energy with the surroundings.

Some reactions only occur in one direction and are described as irreversible, while other reactions are reversible.

In this investigation you will perform reactions in open and closed systems and consider whether the reactions are reversible or irreversible.

PURPOSE

To perform the following reactions and determine which are reversible and which involve open or closed systems. The reactions are:

- formation of hydrated and dehydrated cobalt(II) chloride
- reaction of iron(III) nitrate and potassium thiocyanate
- burning magnesium
- burning steel wool.

MATERIALS

- small bottle of solid hydrated cobalt chloride
- 0.1 mol L⁻¹ iron(III) nitrate solution, Fe(NO₃)₃
- 0.1 mol L⁻¹ potassium thiocyanate solution, KSCN
- 3–4 cm magnesium ribbon
- steel wool
- 2 semi-micro test-tubes
- semi-micro test-tube rack
- dropping pipette
- white tile
- Bunsen burner
- distilled water
- crucible
- tripod and gauze mat
- lighter
- tongs
- fireproof metal lid
- bench mat
- spatula
- 9V battery

PRE-LAB SAFETY INFORMATION

Materials	Hazard	Control
cobalt chloride	may cause skin, eye and respiratory tract irritation or burns	Wear gloves, lab coat, safety glasses.
Fe(NO ₃) ₃ solution	irritating to the eyes, respiratory system and skin	Wear gloves, lab coat, safety glasses.
KSCN solution	irritating to the skin and eyes	Wear gloves, lab coat, safety glasses.
magnesium ribbon	Burning produces intense white light which can be harmful to eyes; could burn skin and clothing.	Use tongs and eye protection; do not look directly into flame.
iron	Burning iron becomes very hot; could burn.	Use tongs.

Please indicate that you have understood the information in the safety table.

Name (print): _____

I understand the safety information (signature): _____



Part A—Formation of hydrated and dehydrated cobalt(II) chloride

PROCEDURE

- 1 Set up the Bunsen burner, tripod and gauze mat on bench mat with the crucible on top.
- 2 Using the spatula, add a small amount of cobalt chloride to the crucible.
- 3 Light the Bunsen burner and watch the colour changes of the cobalt chloride. Record your observations in Results table 1.
- 4 Allow the crucible to cool and add a little distilled water, and record any colour change in the table.
- 5 Heat the crucible again and observe the colour changes of the cobalt chloride.

.....
PRACTICAL ACTIVITY 5.1

RESULTS

.....

TABLE 1

Time (mins)	Observations
0 minutes	
1 minute	
2 minutes	
5 minutes	
After adding water	
After further heating	

DISCUSSION

.....

- 1 The formula of hydrated cobalt chloride is $\text{CoCl}_2 \cdot 6\text{H}_2\text{O}$ and of dehydrated cobalt chloride is CoCl_2 . Write an equation for this system.

- 2 What colour is the hydrated form of cobalt chloride?

- 3 Describe what occurs when you heated the hydrated cobalt chloride.

- 4 Describe what occurred when you added a little water to the heated cobalt chloride.

Part B—Reaction of iron(III) nitrate and potassium thiocyanate

PROCEDURE

.....

- 1 Using a dropping pipette, add approx. 2 mL of $\text{Fe}(\text{NO}_3)_3$ solution to a semi-micro test tube. Clean the pipette and add approx. 2 mL of KSCN solution to the $\text{Fe}(\text{NO}_3)_3$ solution. The reaction that occurs forms FeSCN^{2+} ions in solution.
- 2 Note the colours of the $\text{Fe}(\text{NO}_3)_3$ solution and the KSCN solution and also the colour of FeSCN^{2+} ions in solution in the Results table. In order to produce reliable and accurate results, view the solutions by looking down the test-tubes so you look through the entire solution.
- 3 Divide the FeSCN^{2+} solution equally into 2 test-tubes.
- 4 Clean the dropping pipette and add 10 drops of $\text{Fe}(\text{NO}_3)_3$ solution to one test-tube. Record the changes.
- 5 Clean the dropping pipette and add 10 drops of KSCN solution to the other test-tube. Record the changes.

RESULTS

.....

TABLE 2

Solution	Colour	Explanation
$\text{Fe}(\text{NO}_3)_3$		no explanation required
KSCN		no explanation required
FeSCN^{2+}		no explanation required
$\text{FeSCN}^{2+} + \text{Fe}(\text{NO}_3)_3(\text{aq})$		
$\text{FeSCN}^{2+} + \text{KSCN}(\text{aq})$		

.....
PRACTICAL ACTIVITY 5.1

DISCUSSION

.....

- 1 Write an equation for this system.

- 2 Explain the colour changes you have observed by completing Results table 2 above.

Part C—Burning magnesium

PROCEDURE

.....

- 1 Set up a Bunsen burner on a bench mat and light the burner.
- 2 Record the colour of the magnesium ribbon.
- 3 Hold one end of the magnesium ribbon with tongs and heat the other end in the flame until it ignites.
- 4 Hold the burning metal at arm's length over a bench mat. Do not look directly at the light produced.
- 5 Record the colour of the powder produced.

RESULTS

.....

Colour of magnesium: _____

Colour of product, magnesium oxide: _____

DISCUSSION

.....

Write the equation for this reaction.

Part D—Burning steel wool

PROCEDURE

.....

- 1 Place the metal lid on the bench mat.
- 2 Make a small ball of steel wool and place it on the metal lid. Have a spatula nearby in case the steel wool becomes entangled with the battery.
- 3 Quickly touch the 9V battery terminals to the steel wool and move it away without allowing it to be caught by the wire wool. Record what you observe.
- 4 Optional: Place a small, non-flammable board (preferably a cement sheet) on top of a digital balance. Tare the balance (reset the scale to zero). Then perform the experiment on the scales and record your observations.

RESULTS

.....

DISCUSSION

.....

- 1 Write the equation for this reaction, given that the formula of iron(III) oxide is Fe_2O_3 .

.....
PRACTICAL ACTIVITY 5.1

2 Explain your observations.

GENERAL DISCUSSION

1 For each of the four reactions, state whether they are reversible or irreversible, and whether they occur in open or closed systems. Explain your answers.

Formation of hydrated and dehydrated cobalt(II) chloride: _____

Reaction of iron(III) nitrate and potassium thiocyanate: _____

Burning magnesium: _____

Burning steel wool: _____

2 Explain the meaning of the term 'dynamic equilibrium'.

CONCLUSION

3 Summarise and explain your results.

RATING MY LEARNING	My understanding improved	Not confident	←	→	Very confident	I answered questions without help	Not confident	←	→	Very confident	I corrected my errors without help	Not confident	←	→	Very confident
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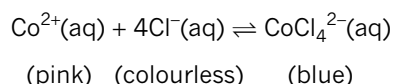
PRACTICAL ACTIVITY 5.2

Investigation of the cobalt chloride equilibrium system

Suggested duration: 20 minutes

INTRODUCTION

Cobalt ions, Co^{2+} , react with chloride ions, Cl^- , to form an equilibrium with the cobalt tetrachloride ion, CoCl_4^{2-} . The reaction is represented by the equation:



If the main species present in an equilibrium mixture is the Co^{2+} ion, then the solution is pink, whereas if the main species present is the CoCl_4^{2-} ion, then the solution is pale blue.

In this experiment you are supplied with a cobalt chloride solution that contains Co^{2+} ions, with lesser amounts of Cl^- and CoCl_4^{2-} ions. The position of the equilibrium (the relative concentrations of ions) in this solution can be changed by altering the temperature, adding more Cl^- ions, or diluting the solution.

PURPOSE

To investigate the position of equilibria established between Co^{2+} , Cl^- and CoCl_4^{2-} ions.

MATERIALS

- 10 mL \times .5 molL⁻¹ cobalt chloride solution CoCl_2
- 4 mL saturated sodium chloride solution NaCl
- 5 mL concentrated hydrochloric acid, HCl
- 10 mL measuring cylinder
- 2 Pyrex[®] test-tubes
- test-tube rack
- stopper
- stirring rod
- Bunsen burner
- bench mat
- 250 mL beaker of ice-water
- safety gloves
- tongs/wooden peg



PRE-LAB SAFETY INFORMATION

Material used	Hazard	Control
concentrated HCl	very corrosive; causes severe burns; harmful by inhalation, ingestion and skin contact	Use small quantities; handle with extreme care and avoid contact; dilute small spills with water. Wear gloves, glasses and a laboratory coat.
$\text{CoCl}_2 \cdot 6\text{H}_2\text{O}$	may cause skin, eye and respiratory tract irritation or burns	Wear gloves, lab coat and safety glasses; wash hands after use.

Please indicate that you have understood the information in the safety table.

Name (print): _____

I understand the safety information (signature): _____

PROCEDURE

Part A—Effect of a change in temperature

- 1 Note the colour of the cobalt chloride solution supplied. Add 5 mL of this solution and 4 mL of saturated sodium chloride solution to a test-tube. Note the colour of the new solution in the Results table 1.
- 2 Briefly heat the solution in the test-tube and note any change in its colour.
- 3 Place the test-tube in a beaker of ice-water. Again note any changes of colour.

Part B—Effect of changes of chloride ion concentration

Slowly and carefully add about 5 mL of concentrated hydrochloric acid to 5 mL of cobalt chloride solution in a test-tube and note the change in colour. Retain this solution for use in Part C.

.....
PRACTICAL ACTIVITY 5.2

Part C—Effect of dilution

To the solution you made in Part B, add sufficient water to double its volume. Stopper and invert the test-tube in order to mix. Note the change in the colour of the solution in Results table 1.

RESULTS

.....

TABLE 1

Part	Solutions	Colour
A	CoCl ₂	
A	CoCl ₂ + NaCl	
A	CoCl ₂ + NaCl after heating	
A	CoCl ₂ + NaCl in ice water	
B	CoCl ₂ + HCl	
C	CoCl ₂ + HCl diluted	

DISCUSSION

.....

- 1 Write an expression for the equilibrium constant, K_{eq} , for this reaction.
.....
- 2 Use your results from Part A to determine how the value of the equilibrium constant changes as the temperature increases. Explain your reasoning.
.....
.....
- 3 Is the reaction exothermic or endothermic?
.....
- 4 Explain why the solution in Part B changes colour when concentrated hydrochloric acid is added.
.....
.....
- 5 Use Le Châtelier's principle to explain the colour change that occurs in Part C when the solution is diluted.
.....
.....

CONCLUSION

.....

- 6 Summarise your results and explanations for the effect of changes to the cobalt chloride equilibrium system.
.....
.....
.....
.....

RATING MY LEARNING	My understanding improved	Not confident	←	→	Very confident	I answered questions without help	Not confident	←	→	Very confident	I corrected my errors without help	Not confident	←	→	Very confident
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PRACTICAL ACTIVITY 5.3

Effect of temperature on equilibrium yields

Suggested duration: 20 minutes

INTRODUCTION

The effect of temperature on an equilibrium depends on whether the reaction is exothermic or endothermic. For this reason, from the way the temperature influences the equilibrium position, you can determine if ΔH is positive or negative. A change in temperature is the only change to an equilibrium system that causes the equilibrium constant to change.

PURPOSE

To investigate the way in which the temperature of an equilibrium mixture influences the position of the equilibrium.

MATERIALS

- 10 mL 1 mol L⁻¹ phosphoric acid H₃PO₄
- methyl violet indicator
- 250 mL beaker of ice-water
- 2 semi-micro test-tubes
- semi-micro test-tube rack
- semi-micro test-tube holder
- 250 mL beaker
- Bunsen burner, gauze mat and tripod stand
- bench mat
- safety gloves
- tongs/wooden peg



PRE-LAB SAFETY INFORMATION

Material	Hazard	Control
1 mol L ⁻¹ phosphoric acid	irritating to the eyes and skin	Wear gloves, safety glasses and lab coat.
methyl violet indicator	may irritate the eyes and skin	Wear gloves, safety glasses and lab coat.

Please indicate that you have understood the information in the safety table.

Name (print): _____

I understand the safety information (signature): _____

PROCEDURE

For the exothermic equilibrium system: $\text{H}_3\text{PO}_4(\text{aq}) \rightleftharpoons \text{H}_2\text{PO}_4^-(\text{aq}) + \text{H}^+(\text{aq})$

- 1 Pour 1 mol L⁻¹ phosphoric acid into each of two semi-micro test-tubes, to a depth of about 3 cm. Add two drops of methyl violet indicator to each test-tube.
- 2 Place one test-tube in ice-water. Carefully heat the other test-tube until the solution almost boils. Record the colour of the indicator in the heated and cooled test-tubes in the Results table.

RESULTS

Colour when cold	Colour when hot

.....
PRACTICAL ACTIVITY 5.3

DISCUSSION

.....

1 In this experiment the colour of the equilibrium mixture is an indication of the concentration of the H^+ ions in solution, $[H^+]$. Methyl violet indicator is yellow in solutions with a high $[H^+]$. Its colour changes through green to blue and then to violet as the concentration decreases. How does $[H^+]$ in the equilibrium mixture change as the temperature is increased?

2 What does this indicate about the value of the equilibrium constant, K_{eq} , for this reaction as the temperature increases? Explain.

3 Describe how the position of equilibrium has shifted and account for your observations in this experiment.

CONCLUSION

.....

4 Summarise the effects of changes in temperature on an equilibrium system.

RATING MY LEARNING	My understanding improved	Not confident	←	→	Very confident	I answered questions without help	Not confident	←	→	Very confident	I corrected my errors without help	Not confident	←	→	Very confident
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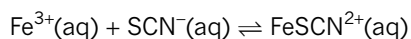
PRACTICAL ACTIVITY 5.4

Determination of K_{eq} by colorimetry

Suggested duration: 1 hour 15 minutes (45 minutes for the experiment and 30 minutes for data analysis)

INTRODUCTION

In aqueous solution, Fe^{3+} ions react with SCN^{-} ions to form the blood-red coloured $FeSCN^{2+}$ ion. The equilibrium is represented by the equation:



The colour of the solution is directly proportional to the concentration of $FeSCN^{2+}$ ions present.

Later in the year, in Module 8, you will learn about the analytical technique called colorimetry. This spectroscopic technique is used to determine the concentration of coloured species in an aqueous sample. The absorption of light energy causes electrons in ions or molecules to jump from a lower energy level to a higher one. The amount of light energy absorbed is directly proportional to the concentration of the coloured species.

In this experiment you will use a colorimeter to measure the absorbance of a standard solution with a known concentration of $FeSCN^{2+}$. The concentration of $FeSCN^{2+}$ ions in other equilibrium mixtures can then be found by measuring their absorbance and comparing the absorbance with the absorbance of the standard. Knowing the initial concentration of Fe^{3+} and SCN^{-} in each mixture, the equilibrium concentration of Fe^{3+} and SCN^{-} can be calculated and a value for the equilibrium constant, K_{eq} , for the reaction can be determined.

PURPOSE

To determine the value of the equilibrium constant for the equilibrium system involving Fe^{3+} ions, SCN^{-} ions and $FeSCN^{2+}$ ions by colorimetric analysis.

MATERIALS

- 20 mL \times 0.0020 mol L⁻¹ potassium thiocyanate solution KSCN
- 30 mL \times 0.0020 mol L⁻¹ iron(III) nitrate solution, $Fe(NO_3)_3$
- 15 mL \times 0.0020 mol L⁻¹ iron(III) thiocyanate ion solution
- 15 mL deionised water
- graduated pipette
- pipette filler
- 5 \times 100 mL beakers
- 4 test-tubes
- test-tube rack
- stirring rod
- colorimeter and data collection device
- colorimeter cell
- marking pen
- tissues
- safety gloves



PRE-LAB SAFETY INFORMATION

Material	Hazard	Control
KSCN solution	irritating to the skin and eyes	Wear gloves, lab coat, safety glasses.
$FeSCN^{2+}$ solution	contains nitric acid and is irritating to the skin and eyes	Wear gloves, lab coat, safety glasses.
$Fe(NO_3)_3$ solution	irritating to eyes, respiratory system and skin	Wear gloves, lab coat and safety glasses.

Please indicate that you have understood the information in the safety table.

Name (print): _____

I understand the safety information (signature): _____

PROCEDURE

- 1 Label four test-tubes A to D.
- 2 Using a graduated pipette, place volumes of 0.0020 mol L⁻¹ $Fe(NO_3)_3$ solution, 0.0020 mol L⁻¹ KSCN solution and water into each test-tube as indicated in the table below. Mix each solution thoroughly with a stirring rod. Be careful to clean and dry the stirring rod after each mixing.

PRACTICAL ACTIVITY 5.4

Test-tube	Fe(NO ₃) ₃ (aq) (mL)	KSCN(aq) (mL)	H ₂ O (mL)
A	5.0	2.0	3.0
B	5.0	3.0	2.0
C	5.0	4.0	1.0
D	5.0	5.0	0.0

- Prepare a reference cell or 'blank' by filling the colorimeter cell to three-quarters of its volume with deionised water. Wipe the outside of the cell with a tissue. By following the manufacturer's instructions, calibrate the colorimeter to read zero transmittance when no light passes through the cell and 100% transmittance when blue light (470 nm) passes through the blank. Use blue light for the remainder of this experiment.
- Discard the deionised water from the cell, rinse the cell twice with the standard 0.00020 mol L⁻¹ FeSCN²⁺ solution and fill the cell to three-quarters of its volume. Measure the absorbance and record the results in the Results table.
- Repeat the procedure in step 4 to measure the absorbance of the solutions in each of test-tubes A–D.

RESULTS

Solution	Absorbance
FeSCN ²⁺	
test-tube A	
test-tube B	
test-tube C	
test-tube D	

DISCUSSION

- Write an expression for the equilibrium law for the reaction.

- Use the following table to enter the results of the calculations in questions 3–8 below.

	Test-tube			
	A	B	C	D
[Fe ³⁺] _{initial}				
[SCN ⁻] _{initial}				
[FeSCN ²⁺] _{equilibrium}				
[Fe ³⁺] _{equilibrium}				
[SCN ⁻] _{equilibrium}				
K _{eq}				

.....
PRACTICAL ACTIVITY 5.4

- 3 Calculate the initial concentration of Fe^{3+} in each of the test-tubes A–D and fill it in the appropriate space in the table. The $[\text{Fe}^{3+}]$ in each tube is given by the expression:

$$[\text{Fe}^{3+}]_{\text{initial}} = \frac{0.0020 \times \text{volume of Fe}(\text{Nin}_3)_3 \text{ the tube in mL}}{\text{total volume in mL}}$$

- 4 Calculate the initial concentration of SCN^- in each of test-tubes A–D and put this in the table. $[\text{SCN}^-]$ in each tube is given by the expression:

$$[\text{SCN}^-]_{\text{initial}} = \frac{0.0020 \times \text{volume of KSCN in the tube in mL}}{\text{total volume in mL}}$$

- 5 The absorbance of light by a solution is directly proportional to the concentration of the absorbing chemical in solution. As a result, you can calculate the concentration of FeSCN^{2+} at equilibrium in each test-tube by comparing the absorbance of each solution with the absorbance of the standard solution.

$$[\text{FeSCN}^{2-}]_{\text{equilibrium}} = 0.00020 \times \frac{\text{absorbance of solution}}{\text{absorbance of standard}}$$

- 6 Calculate the $[\text{Fe}^{3+}]_{\text{equilibrium}}$ in each of the tubes:

$$[\text{Fe}^{3+}]_{\text{equilibrium}} = [\text{Fe}^{3+}]_{\text{initial}} - [\text{FeSCN}^{2+}]_{\text{equilibrium}}$$

- 7 Calculate the $[\text{SCN}^-]_{\text{equilibrium}}$ in each of the tubes:

$$[\text{SCN}^-]_{\text{equilibrium}} = [\text{SCN}^-]_{\text{initial}} - [\text{FeSCN}^{2+}]_{\text{equilibrium}}$$

- 8 Using the equilibrium law from question 1, calculate a value for the equilibrium constant, K_{eq} , for each set of data.

.....
CONCLUSION

- 9 Calculate an average value for K_{eq} . Comment on the reliability and accuracy of K_{eq} determined from each set of data.

RATING MY LEARNING	My understanding improved	Not confident	←	→	Very confident	I answered questions without help	Not confident	←	→	Very confident	I corrected my errors without help	Not confident	←	→	Very confident
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PRACTICAL ACTIVITY 5.5

Precipitation reactions

Suggested duration: 50 minutes

INTRODUCTION

When two solutions containing dissolved ionic salts are mixed together, an insoluble product called a precipitate may form and settle out of the mixture. Knowing which ions form precipitates is essential in many industrial processes, and in monitoring and maintaining the health of natural waterways.

PURPOSE

Part A: To observe reactions that involve the formation of a precipitate.

Part B: To distinguish between ions that form precipitates and those that are always soluble.

To write net ionic equations to represent the formation of precipitates.

PRE-LAB SAFETY INFORMATION

Material	Hazard	Control
sodium hydroxide	corrosive to skin and eyes	Wear eye and skin protection.
sodium carbonate	slightly toxic if ingested	Wear eye and skin protection.
calcium nitrate	slightly toxic if ingested	Wear eye and skin protection.
copper(II) nitrate	slightly toxic if ingested and by skin contact	Wear eye and skin protection.
zinc nitrate	slightly toxic if ingested	Wear eye and skin protection.
silver nitrate	toxic if ingested, stains skin black	Wear eye and skin protection.
barium nitrate	toxic if ingested	Wear eye and skin protection.
potassium nitrate	may irritate eyes and skin	Wear eye and skin protection.

Please indicate that you have understood the information in the safety table.

Name (print): _____

I understand the safety information (signature): _____

Part A—Solubility of ionic compounds

PROCEDURE

- Using a marker pen, write the anions you are testing along the top of the plastic well tray. Write the cations you are testing down the side of the wells. (Results table 1 below shows how your grid should look.) Place the well tray on a dark background for easier observation.
- Place two drops of the appropriate cation solution and one drop of the appropriate anion solution into the wells, according to the labels you have written. Be careful not to allow any contamination, in order to produce reliable and accurate results.
- If no precipitate is formed, record 's' for soluble.
- If a precipitate forms, record 'ppt' (for precipitate) and record its colour.

MATERIALS

- dropper bottles each containing 10 mL of the following 0.5 mol L⁻¹ solutions
 - sodium sulfate Na₂SO₄
 - sodium chloride NaCl
 - sodium hydroxide, NaOH
 - sodium carbonate Na₂CO₃
- dropper bottles each containing 10 mL of the following 0.1 mol L⁻¹ solutions
 - magnesium nitrate Mg(NO₃)₂
 - calcium nitrate Ca(NO₃)₂
 - copper(II) nitrate Cu(NO₃)₂
 - zinc nitrate Zn(NO₃)₂
 - potassium nitrate KNO₃
 - silver nitrate AgNO₃
 - barium nitrate Ba(NO₃)₂
- plastic well tray or plastic grid
- marker pen
- 5 dropper bottles randomly labelled A–, containing 10 mL of the following 0.1 mol L⁻¹ solutions
 - potassium nitrate KNO₃
 - sodium chloride NaCl
 - sodium sulfate Na₂SO₄
 - sodium hydroxide, NaOH
 - sodium carbonate Na₂CO₃
- safety gloves



.....
PRACTICAL ACTIVITY 5.5

RESULTS

Ion	Nitrate (NO ₃ ⁻)	Chloride (Cl ⁻)	Sulfate (SO ₄ ²⁻)	Hydroxide (OH ⁻)	Carbonate (CO ₃ ²⁻)
Mg ²⁺					
K ⁺					
Ca ²⁺					
Cu ²⁺					
Zn ²⁺					
Ag ⁺					
Ba ²⁺					

DISCUSSION

1 What generalisations can you make about the solubilities of:

a nitrates?

b chlorides?

c hydroxides?

d sulfates?

e carbonates?

2 Which cation(s) did not form any precipitates when combined with the anions?

3 Write a balanced ionic equation to represent the formation of each precipitate observed.

PART B—Identifying unknown solutions

PROCEDURE

.....

You have been given one of five unknown solutions in dropper bottles. Your dropper bottle may contain a solution of potassium nitrate (KNO_3), sodium chloride (NaCl), sodium sulfate (Na_2SO_4), sodium hydroxide (NaOH) or sodium carbonate (Na_2CO_3).

- 1 Record the label on your unknown sample: _____
- 2 Select your own set of reagents that will enable you to identify the given unknown sample.
- 3 Record your results in the Results table below.

RESULTS

.....

Reagent	Observations

DISCUSSION

.....

- 1 Write the chemical formula for each precipitate formed in your investigation.

- 2 Identify your unknown sample. Support your answer with experimental evidence.

- 3 It is often possible to convert one ionic compound to another by exploiting solubility differences. Briefly describe, in a series of steps, how you could convert calcium chloride to potassium chloride.

.....
PRACTICAL ACTIVITY 5.5

CONCLUSION
.....

4 Summarise your results in this experiment by writing a series of solubility rules based on your observations.

5 Explain the benefits of writing an ionic equation for each of these reactions.

RATING MY LEARNING	My understanding improved	Not confident	←	→	Very confident	I answered questions without help	Not confident	←	→	Very confident	I corrected my errors without help	Not confident	←	→	Very confident
		<input type="radio"/>		<input type="radio"/>	<input type="radio"/>		<input type="radio"/>		<input type="radio"/>	<input type="radio"/>		<input type="radio"/>		<input type="radio"/>	<input type="radio"/>

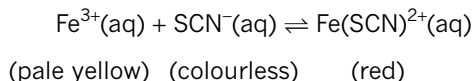
DEPTH STUDY 5.1

Practical investigation of an equilibrium system

Suggested duration: 2 hours 15 minutes (including writing time)

INTRODUCTION

In this activity you will use a solution of $\text{Fe}(\text{SCN})^{2+}$ that contains the ions Fe^{3+} , SCN^- and $\text{Fe}(\text{SCN})^{2+}$ at equilibrium:



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

If, for example, the colour of the solution deepens, the amount of $\text{Fe}(\text{SCN})^{2+}$ ions has increased and the amount of the Fe^{3+} and SCN^- ions must have simultaneously decreased because these ions are used up to form more $\text{Fe}(\text{SCN})^{2+}$. The equilibrium would be described as having a net forward reaction (its position would have 'shifted to the right').

In Part A you will investigate the effect of concentration changes on an aqueous equilibrium.

In Part B you will investigate the way in which the temperature of an equilibrium mixture influences the position of the equilibrium.

QUESTIONING AND PREDICTING

In this investigation you are using an equilibrium system containing the ions Fe^{3+} , SCN^- and $\text{Fe}(\text{SCN})^{2+}$. Use your knowledge of equilibrium and the information in the Introduction above to develop a hypothesis for the outcomes of Parts A and B of this practical investigation.

PRE-LAB SAFETY INFORMATION

As directed by your teacher, complete the risk assessment and management table by referring to the hazard labels on the reagent bottles or safety data sheets (SDS) or your teacher's risk assessment for the activity.



MATERIALS

- 35 mL of $5 \times 10^{-4} \text{ mol L}^{-1}$ iron(III) thiocyanate ion solution $\text{Fe}(\text{SCN})^{2+}$ (20 mL of 0.1 mol L^{-1} $\text{Fe}(\text{NO}_3)_3$ and 20 mL of 0.1 mol L^{-1} potassium thiocyanate (KSCN) per litre)
- 0.1 mol L^{-1} iron(III) nitrate solution, $\text{Fe}(\text{NO}_3)_3$
- 0.1 mol L^{-1} potassium thiocyanate solution, KSCN
- 0.1 mol L^{-1} sodium fluoride solution NaF
- 0.1 mol L^{-1} silver nitrate solution AgNO_3
- 6 semi-micro test-tubes
- semi-micro test-tube holder
- semi-micro test-tube rack
- 250 mL beaker of ice-water
- 250 mL beaker
- Bunsen burner gauze mat and tripod stand
- bench mat
- marking pen
- dropping pipette
- white tile or white sheet of paper
- safety gloves

Material	Hazard	Control
0.1 mol L^{-1} silver nitrate solution	can stain skin, clothing and bench surfaces	
0.1 mol L^{-1} iron(III) nitrate solution		

Please indicate that you have read and understood the information in the table above.

Name (print): _____

I understand the safety information (signature): _____

Part A—Effect of concentration changes on equilibrium yields

CONDUCTING YOUR INVESTIGATION

In this part of the investigation you will perform five tests, labelled A to E, to study the effect on the equilibrium system when the amount (number of moles) of Fe^{3+} or SCN^- ions present in the solution is initially changed as follows.

- In Test A, $\text{Fe}(\text{NO}_3)_3$ is added, which increases the amount of Fe^{3+} .
- In Test B, KSCN is added, which increases the amount of SCN^- .
- In Test C, NaF is added, which decreases the amount of Fe^{3+} because F^- ions react with Fe^{3+} ions to form $\text{FeF}_6^{3-}(\text{aq})$.
- In Test D, AgNO_3 is added, which decreases the amount of SCN^- because Ag^+ ions react with SCN^- ions to form a white precipitate of AgSCN .
- In Test E you will study the effect of the addition of water (dilution) on the equilibrium.

PROCEDURE

- 1 Fill each of six semi-micro test-tubes to one-third of its volume with $\text{Fe}(\text{SCN})^{2+}$ 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' to 'F'.
- 2 Using test-tube F for comparison, perform each of the tests described in Results table 1 and record the change that occurs in the colour of the solution when viewed down the test-tube.

RESULTS TABLE 1 Colour changes

Test-tube	Test	Colour change
A	1 drop of $\text{Fe}(\text{NO}_3)_3(\text{aq})$ added	
B	1 drop of $\text{KSCN}(\text{aq})$ added	
C	1 drop of $\text{NaF}(\text{aq})$ added	
D	1 drop of $\text{AgNO}_3(\text{aq})$ added	
E	Equal volume of water added	
F	None	No change

DISCUSSION

- 1 Write an expression for the equilibrium constant of the reaction that is the subject of this practical investigation.
- 2 Complete Table 2 below for each test by stating:
 - a the initial effect on the concentration of Fe^{3+} or SCN^- of each test
 - b the concentration change of $\text{Fe}(\text{SCN})^{2+}$ after the test
 - c the direction in which the position of equilibrium has shifted.

The entries for Test A have been completed for you.

TABLE 2 Equilibrium results

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

.....

DEPTH STUDY 5.1

3 Sketch concentration–time graphs to show how the concentration of each ion has changed during each test.

4 Sketch rate–time graphs to show how the rates of the forward and reverse reactions have changed during each test.

.....
DEPTH STUDY 5.1

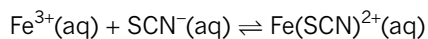
5 Use Le Châtelier's principle to account for the way in which the position of equilibrium shifts in Test A.

6 Account for the way in which the position of equilibrium shifts in Test E, using Le Châtelier's principle.

7 Use collision theory to explain the way in which the position of equilibrium shifts in Test B.

Part B—Effect of temperature on equilibrium yields

1 Design an investigation that will allow you to determine the effect of a change in temperature on the following equilibrium system:



Write your method in the space below.

2 From the list of materials provided above, select and record the ones you will require for your investigation.

3 Check your method and materials with your teacher and then carry out your investigation. Record your results below in a suitable table.

.....

DEPTH STUDY 5.1

Sketch one or more concentration-time graphs to show how the concentration of each ion has changed during your investigation.

Use Le Châtelier's principle and the collision theory to explain the effects of a change in temperature on the equilibrium system. In your explanation, state how the value of the equilibrium constant, K_{eq} , for this reaction changes as the temperature changes and whether the reaction is endothermic or exothermic.

COMMUNICATING YOUR FINDINGS AND PRESENTING THE REPORT

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Write a practical report for this investigation as a Word document or similar. The questions you have answered above and the sections below will guide your report, but you can also add further to your discussion. You can refer to the Toolkit on page ix for what should be included in each section of a practical report. Don't forget to include a conclusion and references.

Processing data and information

Devise appropriate tables and graphs for inclusion in your final report that record all of the raw data you collected in Parts A and B of the investigation.

Analysing data and information

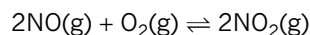
Devise tables, graphs or flow charts as appropriate in which you can make generalisations about the effects of the changes to the equilibrium system. You should add additional analyses of your results and/or the investigation as you see fit.

Discuss any errors or limitations in the data you have collected in this investigation. What modifications, if any, can you make to your hypothesis based on the new evidence collected in this investigation? You should consider the accuracy and reliability of your results and suggest improvements that could be made to this investigation.

Multiple choice

- 1 Which one of the following best describes a chemical system that is in equilibrium?
- A The rate of conversion of reactants to products is negligible.
- B All reaction has ceased.
- C The concentrations of the reactants are equal to those of the products.
- D The rates of the forward and reverse reactions are equal.

- 2 NO reacts with O₂ to form NO₂ according to the equation:

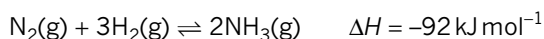


When equilibrium is established at a fixed temperature, which one of the following fractions would be a constant, irrespective of the initial concentrations of the gases?

- A $\frac{[\text{NO}_2]^2}{[\text{NO}]^2[\text{O}_2]}$
- B $\frac{[\text{NO}_2]}{[\text{NO}][\frac{1}{2}\text{O}_2]}$
- C $\frac{[\text{NO}_2]}{[\text{NO}][\text{O}_2]}$
- D $\frac{[\text{NO}_2]}{[\text{NO}][\text{O}_2]^2}$
- 3 Which one or more factors affect the value of the equilibrium constant, K_{eq} , of a gaseous equilibrium?
- A total number of particles in the system
- B total pressure of the system
- C volume of reaction vessel
- D temperature of the system

The following information refers to questions 4 and 5.

N₂ forms an equilibrium with H₂ and NH₃ according to the equation:

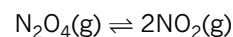


- 4 Which one of the following occurs when the volume of an equilibrium mixture of these gases is reduced at constant temperature?
- A The value of K_{eq} of the system increases.
- B The rate of the forward reaction decreases and the rate of the reverse reactions increases
- C The mass of NH₃ increases and the mass of N₂ and H₂ decreases.
- D The total number of gas particles increases.

- 5 Which one of the following occurs when the temperature of an equilibrium mixture of these gases is decreased?
- A The value of K_{eq} of the system increases.
- B The rate of the forward reaction decreases and the rate of the reverse reactions increases.
- C The mass of NH₃ decreases and the mass of N₂ and H₂ increases.
- D The total number of gas particles increases.
- 6 If the solubility of magnesium hydroxide, Mg(OH)₂, is represented as $s \text{ mol L}^{-1}$, which one of the following expressions is equal to the solubility product, K_{sp} , of magnesium hydroxide?
- A s
- B s^2
- C $3s^2$
- D $4s^3$

Short answer

- 7 N₂O₄ forms an equilibrium with NO₂:



The equilibrium constant is 0.72 at 250°C. If an equilibrium mixture contains 0.020 mol L⁻¹ N₂O₄ at 250°C, calculate the equilibrium concentration of NO₂ in the mixture.

- 8 A precipitate will form when the following solutions are mixed. Write a balanced ionic equation for each reaction.
- a silver nitrate and calcium chloride
-
- b lead(II) nitrate and sodium carbonate
-
- c copper(II) sulfate and ammonium phosphate
-
- d aluminium sulfate and potassium hydroxide
-
- e iron(III) nitrate and sodium sulfide
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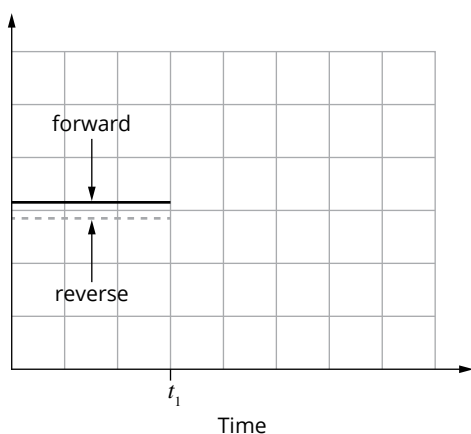
- 9 Consider the following equation that represents a gas-phase reaction in a state of equilibrium:



A scientist performed a number of experiments to investigate the effect of changes on the equilibrium.

a In one experiment, after time, t_1 , more reactant A was added rapidly to an equilibrium mixture at a constant temperature.

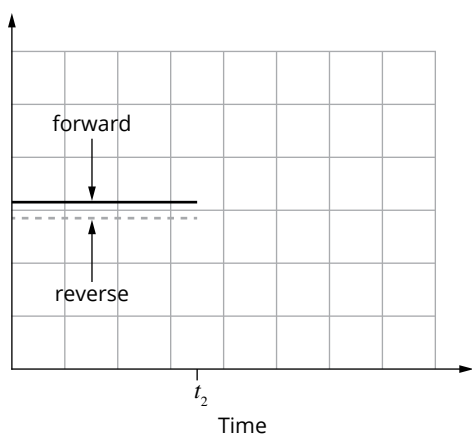
i Complete the graph below to show the effect of this change on the rates of the forward and reverse reactions as the system returned to equilibrium.



ii Explain what happened to the rates of the forward and reverse reactions in this experiment in terms of collision theory.

b In another experiment, at time t_2 a catalyst was added to an equilibrium mixture at a constant temperature.

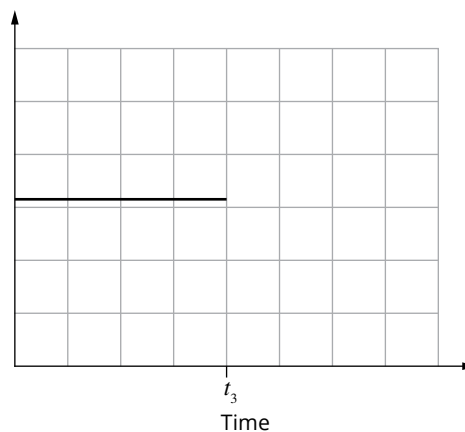
i Complete the graph below to show the effect of the catalyst on the rates of the forward and reverse reaction.



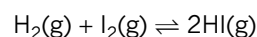
ii Explain, in terms of collision theory, what happened to the rates of the forward and reverse reactions in this experiment.

c In a third experiment, after time t_3 the temperature was rapidly increased to a new, constant value and equilibrium was re-established.

Complete the concentration–time graph below to show the effect of the temperature change on the concentration of C.



d 2.00 mol of HI was placed in a 1.00L container at 600K. The HI decomposed, forming 0.22 mol of H_2 and 0.22 mol of I_2 at equilibrium. Calculate the value of the equilibrium constant, K_{eq} , for the reaction:



at this temperature.

10 A yellow precipitate of lead iodide forms when a solution of lead nitrate is mixed with a solution of potassium iodide.

a Name the spectator ions in this reaction.

b Write a balanced ionic equation for the formation of the precipitate.

c Write an expression for the solubility product, K_{sp} , of the yellow solid that is formed in this reaction.

d At 25°C, 1.00L of water can dissolve 1.26×10^{-3} mol of lead iodide. Calculate the solubility product, K_{sp} , for lead iodide at this temperature.
