**Chemistry Module 5: Equilibrium and Acid Reactions**

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## Course overview

The chemistry course develops student ability to apply processes that underpin the understanding of the nature and practice of science. It enables students to interpret the interconnection between the nature and practice of science and knowledge of chemistry. Through applying the Working Scientifically skills processes, the course aims to examine how theories, models and practices are used and developed.

**The fundamentals developed in Year 11 include**:

* knowledge, understanding of the properties and structures of matter, and relating properties and structures to trends in data and reactions.
* knowledge and understanding of the types of and drivers of chemical reactions.
* skills in making hypotheses and designing valid and reliable practical investigations.
* skills in conducting investigations and solving problems.
* constructing models.
* conducting investigations by measuring and/or collecting relevant data and information from first-hand practicals and secondary sources and determining the accuracy, reliability and validity of data and information.

**The Year 12 course builds on the concepts introduced in Year 11 by:**

* examining classes of chemicals, processes and a variety of chemical reactions which incorporate organic compounds and acid/base equilibrium reactions
* the application of this knowledge to the investigation of a range of methods used in identifying and measuring quantities of chemicals, which leads to a deep understanding of the structure, properties and trends of, and between, classes of chemicals
* using physical, conceptual and mathematical models to predict outcomes and assessing the limitations of models.

## Module summary

This module addresses the concepts of equilibrium in both qualitative and quantitative approaches for students to explore the nature of chemical reactions. Students use Le Chatelier’s principle to predict and analyse reactions, with reference to conceptual models, analogies, graphical representations and mathematical models. Students link thermodynamics and energy to equilibrium and the analysis of reactions, including solubility and dissolution.

## Big ideas

* Chemical and physical changes may be described in open and closed systems.
* Static and dynamic equilibrium may be modelled.
* There may be no observable changes at the macro level, but at a molecular level, equilibrium can be observed, modelled, and described.
* A disturbance to a system will cause the equilibrium position to shift and these changes can be predicted, described and modelled using Le Chatelier’s principle and collision theory.
* Equilibrium can be analysed using various representations of qualitative and quantitative data.
* Changes in energy and the analysis of energy profile diagrams are linked to changes in the equilibrium position.
* Mathematical models can be used to predict, explain and analyse a range of reactions, including solubility and dissociation of ionic compounds and acids.

## Relationship to other modules

Learning about reactions, equations and stoichiometry in these Year 11 modules will provide the basis for the work on quantitative equilibrium:

* Module 2 Introduction to Quantitative Chemistry, (IQ2-1 Chemical Reactions and Stoichiometry and IQ2-3 Concentration and Molarity and IQ2-4 Gas Laws)
* Module 3 Reactive Chemistry, (IQ3-1 Chemical Reactions, IQ3-3 Rates of Reaction).

Factors that drive reactions will be built on knowledge and understanding from:

* Module 4 Drivers of reactions, (IQ4-1 Energy Changes in Chemical Reactions and IQ4-3 Entropy and Gibbs Free Energy)

Module 5 is best taught in the order in which it appears in the syllabus. If Module 6 is taught before Module 5 the concept of Ka may be difficult to describe without the first two inquiry questions from Module 5 being covered first.

## Core concepts

* Models are constructed, analysed and assessed as analogues for chemical reactions.
* Models of equilibrium can be qualitative or quantitative, and can be represented as mathematical models, graphical models, simulations and animations, conceptual models and physical models.
* Practical investigations are designed and conducted to analyse reactions using a range of data representations, including constructing balanced equations with states and enthalpy of reaction.
* Graphs and energy profile diagrams are constructed and analysed.
* Factors and changes to reactions are predicted and analysed qualitatively by applying Le Chatelier’s principle and quantitatively through the calculation of equilibrium constant, K.
* By constructing ICE tables, equilibrium constants (K) and reaction quotients (Q) are calculated, using data from a range of representations.
* Knowledge of equilibrium is applied to solubility and dissociation of ionic compounds and to dissociation of acids and bases.
* The calculating of equilibrium constants is the basis of the ability to analyse and predict reactions.

## Opportunities for extending concepts

Extension of concepts may be used as a starting point for depth studies. For example:

* Links to entropy and enthalpy as drivers of reactions
* Predicting reactions, reaction direction and spontaneity

## Alternative conceptions and misconceptions

* Chemical reactions are irreversible.
* Reactions that are proceeding must exhibit observable changes and if no changes are evident the reaction is stopped.
* Rates of forward and reverse reactions are not equal at equilibrium.
* Concentrations of products and reactants at equilibrium are equal.
* Rate of reaction and equilibrium are the same thing, that is, if the rate is faster, more product will be made.
* Animations, simulations and models may assist students to visualise the understand equilibrium, collision theory and rates of reaction at a particle level.
* Models are always physical 2-D and 3-D representations.
* Mathematical models are also able to summarise equilibrium, as evidenced by calculation and analysis of reaction quotient, Q and equilibrium constant, K.

## Conceptual difficulties

* Applying models and analogies to reactions and collision theory.
* The difference between static and dynamic equilibrium and [open and closed systems](https://youtu.be/JsoawKguU6A). [This video](https://www.flinnsci.com/equilibrium-demonstrations---the-good-the-bad-and-the-ugly2/vel1735/) discusses the conceptual difficulties in understanding changes to an equilibrium system vs reactants and products sides.
* Using qualitative and quantitative data to analyse equilibrium, understanding which to use and predicting the effect of changes on equilibrium. A thorough understanding of the nature of qualitative equilibrium is necessary before students analyse equilibrium quantitatively.
* Assessing the [validity and reliability of data](https://schoolsequella.det.nsw.edu.au/file/ee66cc99-c090-42d7-8bc8-85734c19a0b9/1/Evaluating_data.docx). Practical activities provide opportunities for students to design investigations and analyse the data collected.
* Optimising a process by considering both rate and the factors that affect equilibrium. Modelling, animations and simulations may be used to evaluate a process, once students are competent in visualising and understanding equilibrium and rate of reaction.
* Mathematical concepts involving:
	+ the writing of expressions for equilibrium constant, K
	+ the calculation and assumptions in determining the concentration of various species
	+ coefficients in stoichiometry and indices
	+ uncertainty and assumptions.

Assigning problem sets and then applying formative assessment in the classroom is recommended to assess the mastery of mathematical concepts. Confident students may be extended with more advanced calculations and the analysis of equilibrium reactions.

## Suggested teaching strategies

This is not a teaching program, but a sequence of strategies. This document attempts to address the conceptual difficulties and alternative conceptions as well as the parts of the syllabus that are new or that may be difficult for which to construct effective learning sequences.

### IQ5-1: What happens when chemical reactions do not go through to completion?

To engage students, equilibrium may be introduced with a practical example, with an observable change, such as the cobalt chloride equilibrium. Teachers should complete a risk assessment prior to choosing which equilibrium reaction will be used, or alternatively use media recordings of observable reactions. Chemical experiments which demonstrate equilibrium systems and introduce the concept of changes to conditions impacting reactions include:

* [Cobalt Chloride (see appendix 2).](https://edu.rsc.org/resources/the-equilibrium-between-two-coloured-cobalt-species/1.article)
* Equilibrium between [nitrogen dioxide and dinitrogen tetroxide](https://edu.rsc.org/resources/le-chateliers-principle-the-equilibrium-between-nitrogen-dioxide-and-dinitrogen-tetroxide/1739.article) (note, copper metal and concentrated nitric acid is a suitable substitute for unsafe heating of lead nitrate to generate nitrogen dioxide gas). The weakness in the practical is the demonstration of pressure changes on nitrogen dioxide, which is difficult to achieve in the laboratory and can be hazardous if precautions are not completed appropriately (see appendix 2).
* [Potassium dichromate](https://edu.rsc.org/resources/an-equilibrium-involving-chromate-vi-and-dichromate-vi-ions/1710.article) are chemical experiences/demonstrations showing qualitative observations of equilibria with colour changes and linking the observations to equations and the practice of core skills for analysis of equilibrium (see appendix 2).

The link with a range of effective non-chemical analogies will assist students to qualitatively describe chemical equilibrium. The factors which impact dynamic equilibrium can be linked to the chemical equations used. Large scale models with students as molecules can be very engaging.

### IQ5-2: What factors affect equilibrium and how?

Le Chatelier’s principle is well described in many texts and online resources. The application of Le Chatelier’s principle to a range of reactions and representations, including models, will demonstrate deep understanding.

The introduction of graphs, representing changes to a system and the effect on species concentrations, can represent a large amount of information in an organised, visual way. Students can describe changes, conditions or concentrations by referring to Le Chatelier’s principle and both the equation and the graph. The use of graphical representations of changes to systems as stimulus for formative assessment will enable teachers to assess understanding.

Complete a case study of [Haber process](https://www.learner.org/series/chemistry-challenges-and-solutions/control-a-haber-bosch-ammonia-plant/) (or an alternative [more mathematical simulator](https://demonstrations.wolfram.com/ChemicalEquilibriumInTheHaberProcess/)) as an industrial example of equilibrium in action (links well with IQ8-3), by determining gas moles, balancing the equation and describing conditions that influence the reaction. The strength in this case study is in the simplicity of the equilibrium reaction equation, as all species being in the gaseous state facilitates easier analysis of pressure, volume and temperature considerations. Weakness exists in the simulation on the cost of production/profits however this is a useful talking point about the need to monitor equilibrium from a commercial perspective. A review of rates of reaction and a discussion about percentage of product yield vs the rate at which product is produced, may help misconceptions about the difference between rate and yield. [Using rhubarb](https://edu.rsc.org/resources/rates-and-rhubarb/745.article) as a safer alternative to the hydrochloric acid and sodium thiosulfate practical, for comparing rate of reaction as concentration and/or surface area is changed, can be included as a review of IQ3-3. A good understanding of both equilibrium and rate is essential to be able to discuss the Haber process adequately.

### IQ5-3: How can the position of equilibrium be described and what does the equilibrium constant represent?

* Calculate K and Q with a strong link back to the analogies and case studies, using sample data, before introducing the ICE table concept to analyse reactions. Given chemical equations and initial concentration data, ICE tables can be used to determine equilibrium values for use in K expressions and Q calculations in order to analyse the equilibrium position. Rearranging of the K and Q expressions to determine concentrations of reaction components is also an important process to complete.
* [ICE tables](https://chem.libretexts.org/Bookshelves/Physical_and_Theoretical_Chemistry_Textbook_Maps/Supplemental_Modules_%28Physical_and_Theoretical_Chemistry%29/Equilibria/Le_Chatelier%27s_Principle/Ice_Tables) this chem.libretexts.org resource gives a step by step introduction into setting up and using ICE table to solve for K or Q. See appendices for examples of how to use ICE tables.

### IQ5-4: How does solubility relate to chemical equilibrium?

* Determine solubility equilibrium position to link Ksp and solubility limit for ionic substances. Gravimetric analysis using precipitation reactions will allow the calculation of solubility limits and Ksp for a range of simple ionic substances. See appendices for examples of how to perform Ksp calculations.

# Appendices

## Appendix 1: Equilibrium concepts

Equilibrium is a difficult concept for students to conceptually process. The use of analogies and models is a highly engaging method of developing student capacity and confidence. Care must be taken in the appropriate selection and use of analogies and models as some are more widely applicable than others. Uncontrolled variables, and inadvertent errors can also impact the outcome of the task and complicate the link to equilibrium for students.

[Physical and Theoretical Chemistry Equilibria](https://chem.libretexts.org/Bookshelves/Physical_and_Theoretical_Chemistry_Textbook_Maps/Supplemental_Modules_%28Physical_and_Theoretical_Chemistry%29/Equilibria) uses physical demonstrations of the concepts in a concrete manner is enhanced in this non chemistry-based equilibrium analogy, using water tank transfers. This analogy features student interaction with the concept and ability to adjust variables to model factors that influence equilibrium. The weakness relates to the inaccurate construction of the apparatus to reduce other inputs. Other models include:

* [Equilibrium concept through interactive analogies](https://uwaterloo.ca/chem13-news-magazine/may-2016/feature/equilibrium-concept-through-interactive-analogies) from the University of Waterloo
* [Escher: The Waterfall and Chemical Equilibrium. An Analogy Derived From Art for Chemistry](https://www.academia.edu/1380183/Escher_The_Waterfall_and_Chemical_Equilibrium._An_Analogy_Derived_From_Art_for_Chemistry). A pdf from academia.edu
* [Equilibrium Demonstrations - The Good, the Bad, and the Ugly!](https://www.flinnsci.com/equilibrium-demonstrations---the-good-the-bad-and-the-ugly2/vel1735/) (duration 27:28)

The below activity is adapted from [An Analogy for Chemical Equilibrium](http://www.umanitoba.ca/outreach/crystal/chemistry.html#chemicalequilibrium) for the procedure and student worksheets which contain a familiar activity utilising drinking straws of differing diameter in measuring cylinders. Fire-polished glass tubing or upside down pipettes could be used as an alternative to straws. Other apparatus can be substituted such as larger containers with different sized cups/beakers (the cup or beaker, must be placed flat on the bottom of the large container and liquid allowed to spill over before transfer), being used to transfer liquids from buckets or large beakers. If using this method, a suitable scale to measure the mass of each container after each transfer would be valuable to add quantitative evidence to the analysis of this analogy. Students must be informed that the system must remain closed to ensure validity. If during the transfers any liquid is lost (through spillage for example) or gained (from sources other than the opposite container) this analogy will fail to accurately demonstrate the principle of dynamic equilibrium.

### Student Activity

#### Materials:

* 2x 25mL graduated cylinders per student group
* 2x small diameter drinking straws per student group
* 2x large diameter drinking straws per student group
* Excel spreadsheet template for students to graph collected data (separate download on the curriculum support website)

#### Procedure

##### Part A:

1. Label a 25mL graduated cylinder ‘A’. Fill it to the 25.0mL mark with water. This is the **reactants**. Label a second 25mL graduated cylinder ‘B’. This is the **products**. This cylinder will begin empty.
2. Obtain 2 straws of the same diameters. Label these straws with ‘A’ and ‘B’. Be sure to keep straw ‘A’ with cylinder ‘A’ and straw ‘B’ with cylinder ‘B’.
3. With a partner, simultaneously lower the straws into each of the graduated cylinders. When the straws reach the bottom of the cylinder each partner will place a finger over the opening of the straw (ensure a tight seal) and then transfer the contents to the opposite graduated cylinder and allow the water to drain completely.
4. Accurately record the volume of water in each of the **reactants** and **products** cylinders on the ‘transfer #1’ space in the excel spreadsheet provided.
5. Repeat steps 3 and 4 until 5 successive transfers result in no further change in volume of either cylinder.
6. Obtain the graph plot from excel of this data in Part A. Create a new version of this excel file for Part B.

**This prescribed procedure is applicable to all students, however if the focus is about teaching and practising the Working Scientifically skills, particularly CH12-1, it is suggested that students move to Part D after Part A.**

##### Part B:

What would the graphs look like if the straw were diameters changed? In groups, repeat the steps in Part A using different diameter straws in the same group. Two large diameter straws, two small diameter straws, one small reactants straw and one large products straw or one large reactants straw and one small reactants straw. Obtain a new excel data and graph for comparison.

##### Part C:

What would the graphs look like if the system was open instead of closed? In groups, repeat the steps in Part A while a third person randomly adds and removes small volumes of liquid from either cylinder “A” and cylinder “B”. Obtain a new excel data and graph for comparison. Continue this activity only until the maximum 100 transfers.

**Sample Questions:**

1. Why is cylinder “B” empty at the start of both Part A and Part B?
2. Why was the transferred volume of ‘B’ to ‘A’ zero for transfer #1?
3. If we started with 10mL in ‘A’ and 15mL in ‘B’, can you predict the equilibrium volumes?
4. What changes are occurring at the molecular level even though there is no apparent change in volume?
5. How do each of the graphs produced in Part A and Part B of this simulation differ?
6. When the system was not closed in Part C how did the graph change?
7. What conclusion about equilibrium can be formed from the comparison between open and closed systems shown in Part C?

##### Part D:

Make a prediction and a hypothesis about what affects the equilibrium, using the model in Part A, and test your hypothesis. Collect data and represent the data in order to make a conclusion.

## Appendix 2: Le Chatelier’s principle

The encouraging of discussion and [argumentation](https://edu.rsc.org/education-research/for-the-sake-of-argument/2000241.article) in class will assist in diagnosing misconceptions and encourage deep thinking about equilibrium, particularly after using animations and simulations to demonstrate principles. The use of animations and simulations can form part of an investigation into equilibrium and students can predict, and test their prediction, through making observations. Animations can also assist in constructing conceptual models about the relationship between collision theory and equilibrium and collision theory and rate of reaction.

Dynamic equilibrium can be impacted by changes which influence the reaction equation.

* Temperature impacts reactions differently depending on whether they are endothermic or exothermic. The addition of heat energy will favour the endothermic reaction where ΔH > 0 (positive). The removal of heat energy will favour the exothermic reaction where ΔH < 0 (negative).
* Pressure (with an inverse relationship to volume) only impacts reactions which have gaseous reactants and/or products. Increasing pressure (decreasing volume) at a constant temperature will favour the reaction which produces the least gas moles. Decreasing pressure (increasing volume) at a constant temperature will favour the reaction which produces the most gas moles. The relationship between pressure and volume can be demonstrated with a syringe or through modelling and [simulations](https://phet.colorado.edu/en/simulation/gas-properties).
* Concentration impacts all equilibrium reactions by the introduction or removal of reactants or products involved in the reaction equation or through the interaction of the added substance with equilibrium components (such as causing precipitation). Adding a reactant will favour the forward reaction and adding a product will favour the reverse reaction. Removal of reactants will favour the reverse reaction and removal of products will favour the forward reaction. [Simulations](https://phet.colorado.edu/en/contributions/view/3858) and animations may assist conceptualisation and [boost learning](https://edu.rsc.org/education-research/animations-boost-learning/3010768.article).

Pure liquids and all solids are not participants in equilibrium reactions and so changing their concentrations is irrelevant.

### Student Activities

The following are suggestions only, and teachers may choose to demonstrate or use videos and simulations, should teachers wish to avoid perceived risks and pursue alternatives.

* [Heating cobalt(II) chloride hydrate solution](https://edu.rsc.org/resources/the-equilibrium-between-two-coloured-cobalt-species/1.article) - will demonstrate the effect of concentration and temperature on the equilibrium system. Hydrochloric acid must be diluted to under 10M (35%) to meet CSIS restrictions for Year 12 student use. Please review the CSIS information for [cobalt chloride](https://online.det.nsw.edu.au/ecmjsp/chemicals/chemicals.cfm?mode=viewchemares&chemalpha=C&chemid=513#skipToContent) and [hydrochloric acid](https://online.det.nsw.edu.au/ecmjsp/chemicals/chemicals.cfm?mode=viewchemares&chemalpha=H&chemid=859#skipToContent) before undertaking this activity with students and use an alternative activity should teachers deem this to be prudent.

$$[Co(H\_{2}O)\_{6} ]\_{(aq)}^{2+}+4Cl\_{(aq)}^{-} \leftrightarrow [CoCl4 ]\_{(aq)}^{2-}+6H\_{2}O\_{(l)}$$

$$[Co(H\_{2}O)\_{6}]\_{(aq)}^{2+} is pink and [CoCl\_{4}]\_{(aq)}^{2-} is blue$$

The forward reaction producing the blue ion is endothermic (ΔH > 0) and applying heat will shift equilibrium right, turning pink solutions blue. Removing heat in an ice bath will shift equilibrium left turning blue solutions pink. Silver nitrate solution can be used to precipitate out silver chloride, shifting equilibrium to replace chloride ions that are lost, shifting equilibrium left and turning blue solutions pink. Please review the CSIS information for [silver nitrate](https://online.det.nsw.edu.au/ecmjsp/chemicals/?mode=viewchemares&chemalpha=S&chemid=1517#skipToContent) before undertaking this activity with students.

* Interaction between [nitrogen dioxide and dinitrogen tetroxide](https://edu.rsc.org/resources/le-chateliers-principle-the-equilibrium-between-nitrogen-dioxide-and-dinitrogen-tetroxide/1739.article) equilibrium will demonstrate the effect of concentration and pressure on the equilibrium system. To avoid the use of heavy metals it is possible to generate nitrogen dioxide in a fume hood using concentrated nitric acid and small amounts of copper metal. Please review the CSIS information for [nitric acid](https://online.det.nsw.edu.au/ecmjsp/chemicals/chemicals.cfm?mode=viewchemares&chemalpha=N&chemid=1166#skipToContent) and [nitrogen dioxide](https://online.det.nsw.edu.au/ecmjsp/chemicals/chemicals.cfm?mode=viewchemares&chemalpha=N&chemid=1175#skipToContent) before undertaking this activity with students.

$$2NO\_{2(g)}\leftrightarrow N\_{2}O\_{4(g)}$$

$$Nitrogen dioxide, NO\_{2(g)} is brown and dinitrogen tetroxide, N\_{2}O\_{4(g)} is colourless.$$

Teachers may prefer to observe [Volume Effect on Equilibrium - LeChatelier's Principle Lab Extension](https://www.youtube.com/watch?v=pnU7ogsgUW8) (duration 0:43) due to risk.

* [Chromate/dichromate equilibrium](https://edu.rsc.org/resources/an-equilibrium-involving-chromate-vi-and-dichromate-vi-ions/1710.article) - equilibrium will demonstrate the effect of concentration on the chemical reaction:

$$2[CrO\_{4}]\_{(aq)}^{2-}+2H\_{(aq)}^{+} \leftrightarrow [Cr\_{2}O\_{7}]\_{(aq)}^{2-}+H\_{2}O\_{(l)}$$

$$Chromate ion, 2[CrO\_{4}]\_{(aq)}^{2-} is yellow and dichromate, [Cr\_{2}O\_{7}]\_{(aq)}^{2-} is orange.$$

2M sodium hydroxide and 2M sulfuric acid can be used to alter the concentration of the hydrogen ion and shifting equilibrium to either side in response. Please review the CSIS information for [potassium chromate](https://online.det.nsw.edu.au/ecmjsp/chemicals/chemicals.cfm?mode=viewchemares&chemalpha=P&chemid=1368#skipToContent), [sodium hydroxide](https://online.det.nsw.edu.au/ecmjsp/chemicals/chemicals.cfm?mode=viewchemares&chemalpha=S&chemid=1568#skipToContent) and [sulfuric acid](https://online.det.nsw.edu.au/ecmjsp/chemicals/chemicals.cfm?mode=viewchemares&chemalpha=S&chemid=1648#skipToContent) before undertaking this activity with students.

## Appendix 3: Keq and Q calculations

### Calculating and applying Keq in equilibrium systems

The equilibrium constant Keq can be used to determine the position of equilibrium. The Keq expression can be written as:

$$K\_{eq}=\frac{\left[products\right]^{mole ratio}}{\left[reactants\right]^{mole ratio}}$$

The equilibrium concentrations of each reactant and product are raised to the power of their mole ratio in the balanced equation. Where multiple reactants and/or products exist, each are multiplied together in the expression raised to their individual mole ratio powers. Concentrations may be expressed as molar concentrations or partial pressures. Pure liquids and solids are excluded from these calculations where present in the reaction equation.

Keq can be written for the Haber Process as:

$$N\_{2\left(g\right)}+3H\_{2\left(g\right)} \leftrightarrow 2NH\_{3\left(g\right)}$$

$$K\_{eq}=\frac{[NH\_{3}]^{2}}{[N\_{2}]^{1}×[H\_{2}]^{3}}$$

To calculate Keq the equilibrium concentrations of each reactant and product must be determined. The use of an ICE (initial, change, equilibrium) table can be very useful in this process. Consider the following example of the Haber process.

Calculate Keq for a sealed (closed system) 3 Litre vessel containing 9 moles of nitrogen gas and 3 moles of hydrogen gas which is then allowed to establish equilibrium. At equilibrium, the container is found to contain 0.3 moles of ammonia gas.

Creating an ICE table from the data given in the question (Initial [NH3] = 0M before the reaction):

|  |  |  |  |
| --- | --- | --- | --- |
|  | $$1N\_{2(g)}$$ | $$3H\_{2(g)}$$ | $$2NH\_{3(g)}$$ |
| Initial | 9 moles in 3L=3M | 3 moles in 3L=3M | 0M |
| Change |  |  |  |
| Equilibrium |  |  | 0.3 moles in 3L= 0.1M |

Through the mole ratio in the balanced reaction equation it is then possible to calculate the other equilibrium concentrations and complete the ICE table:

|  |  |  |  |
| --- | --- | --- | --- |
|  | $$1N\_{2(g)}$$ | $$3H\_{2(g)}$$ | $$2NH\_{3(g)}$$ |
| Initial | 3M | 1M | 0M |
| Change | Change is negative due to consumption.$\frac{-0.1}{2}$ x 1 =$ -0.05M$ | Change is negative due to consumption.$\frac{-0.1}{2}$ x 3 =$ -0.15M$ | Change is positive due to production.$$+0.1M$$ |
| Equilibrium | 3M + -0.05$M$ = $2.95M$ | 1M + -0.15$M$ = $0.85M$ | 0.1M |

Now with all equilibrium concentration values the Keq for this reaction can be determined:

$$K\_{eq}=\frac{[NH\_{3}]^{2}}{\left[N\_{2}\right]^{1} ×\left[H\_{2}\right]^{3}}$$

$$K\_{eq}=\frac{0.1^{2}}{2.95^{1}×0.85^{3}}$$

$$K\_{eq}=\frac{0.01}{1.81}$$

$$K\_{eq}=0.006$$

The small value of Keq shows that this reaction equilibrium favours the left with a much higher concentration of reactants than products in the final equilibrium mixture.

The reverse calculation is also possible to determine the equilibrium concentrations from a given Keq value. Consider the same reaction as the above example with Keq given as 1.01, calculate the equilibrium concentration of ammonia where the equilibrium concentrations of nitrogen is 2.00M and hydrogen is 4.00M. The expression for Keq can be rearranged to determine the concentration of ammonia at equilibrium:

$$K\_{eq}=\frac{[NH\_{3}]^{2}}{\left[N\_{2}\right]^{1}×\left[H\_{2}\right]^{3}}$$

$$K\_{eq}\left(\left[N\_{2}\right]^{1}×\left[H\_{2}\right]^{3}\right)=[NH\_{3}]^{2}$$

$$\sqrt{K\_{eq}\left(\left[N\_{2}\right]^{1}×\left[H\_{2}\right]^{3}\right)}=[NH\_{3}]$$

$$∴[NH\_{3}]=\sqrt{K\_{eq}\left(\left[N\_{2}\right]^{1}×\left[H\_{2}\right]^{3}\right)}$$

The known values can be entered into this new equation and the solution calculated:

$$[NH\_{3}]=\sqrt{K\_{eq}\left(\left[N\_{2}\right]^{1}×\left[H\_{2}\right]^{3}\right)}$$

$$[NH\_{3}]=\sqrt{1.01\left(\left[2.00\right]^{1}×\left[4.00\right]^{3}\right)}$$

$$[NH\_{3}]=\sqrt{1.01\left(128.00\right)}$$

$$[NH\_{3}]=\sqrt{129.28}$$

$$[NH\_{3}]=11.4M (3 s.f.)$$

The larger concentration of the final product ammonia can be assumed from the larger value of Keq given. This simple checking process is valuable for students to validate their mathematical process through the theoretical understanding of Keq.

These questions can be very easily extended with additional steps before or after the Keq calculation to convert between concentration and mass/volume of the reaction components. Additional steps could also be added to show how changes in the reaction conditions will impact Keq and/or the concentrations of the reaction components.

More examples of Keq calculations can be found in numerous textbooks and online [equilibrium constant](https://chem.libretexts.org/Bookshelves/Physical_and_Theoretical_Chemistry_Textbook_Maps/Supplemental_Modules_%28Physical_and_Theoretical_Chemistry%29/Equilibria/Chemical_Equilibria/The_Equilibrium_Constant) as well as more examples of how to make and use [ICE Tables](https://chem.libretexts.org/Bookshelves/Physical_and_Theoretical_Chemistry_Textbook_Maps/Supplemental_Modules_%28Physical_and_Theoretical_Chemistry%29/Equilibria/Le_Chatelier%27s_Principle/Ice_Tables).

Students should apply significant figures only at the final answer stage, and should keep the complete calculator display until the final answer is reported.

### Calculating and applying Q in equilibrium systems

To determine the direction a reaction will proceed it is necessary to calculate Q and compare this value to Keq Q is the ‘reaction quotient’ and is an identical mathematical calculation to Keq. Q shows an ‘in progress’ marker for the equilibrium reaction. A simple analogy can be used to describe the difference between Q and Keq consider a running race, the finish line is equilibrium (Keq) and the runner’s current position is Q.

Consider a reaction in which there are no products at the start of the reaction, the reactants need to turn into products for equilibrium to be established. This is shown by the runner (Q) needing to run forwards to reach the finish line (Keq), similarly, the reaction needs to proceed forward to establish equilibrium. Mathematically this is shown as Q is less than Keq.

At equilibrium the runner (Q) has reached the finish line (Keq) and so no further reaction needs to occur. Mathematically this is shown as Q is equal to Keq, neither the forward nor reverse reaction is favoured, and both occur at the same rate.

In the case where the runner (Q) is beyond the finish line (Keq) it is necessary to run in reverse to reach the finish line. The reaction needs to proceed in reverse to establish equilibrium. Mathematically this is shown as Q is greater than Keq.

Consider the same reaction we have been using, for the reaction conditions used Keq was found to be 0.021, during the reaction the concentrations of each component were measured as:

* Nitrogen = 1.2M
* Hydrogen = 0.80M
* Ammonia = 0.0020M

Determine if this reaction is at equilibrium and which direction the reaction must proceed in order to establish equilibrium for the given Keq.

$$N\_{2\left(g\right)}+3H\_{2\left(g\right)} \leftrightarrow 2NH\_{3\left(g\right)}$$

$$Q=\frac{[NH\_{3}]^{2}}{[N\_{2}]^{1}×[H\_{2}]^{3}}$$

$$Q=\frac{0.0020^{2}}{1.2^{1}×0.80^{3}}$$

$$Q=\frac{0.0000040}{1.2×0.512}$$

$$Q=\frac{0.000004}{0.6144}$$

$$Q=6.5×10^{-6}$$

Keq = 0.021 (given)

This finds that Q is not equal to Keq so this system is not at equilibrium, Q is less than Keq so this reaction needs to proceed forward to establish equilibrium with more reactants being converted into products. More on Q can be found on [Chemistry LibreTexts – Reaction Quotient](https://chem.libretexts.org/Bookshelves/General_Chemistry/Book%3A_Chem1_%28Lower%29/11%3A_Chemical_Equilibrium/11.03%3A_Reaction_Quotient).

## Appendix 4: Solubility and Ksp Calculations

### Calculating and applying Ksp in equilibrium systems

Like Keq the equilibrium constant, Ksp can be used to determine the position of an equilibrium, this time in reference to the solubility of a solute in a given solvent. The Ksp expression can be written as:

$$K\_{sp}=\frac{\left[products\right]^{mole ratio}}{\left[reactants\right]^{mole ratio}}$$

In a solubility equilibrium, the reactants are the pure solids or liquids of the solute and are removed from the Ksp calculation, so the equation becomes:

$$K\_{sp}=\left[aqueous product ions\right]^{mole ratio}$$

The equilibrium concentrations of each aqueous product ions are raised to the power of their mole ratio in the dissolution equation. Each product is multiplied together in the expression raised to their individual mole ratio powers. Concentrations are expressed as molar concentrations. Ksp is determined for a set temperature and pH, altering these conditions will change the solubility equilibrium position and so Ksp will change.

Ksp can be written for the dissolution of lead(II) iodide:

$$PbI\_{2\left(s\right)} \leftrightarrow 1Pb\_{\left(aq\right)}^{2+}+2I\_{\left(aq\right)}^{-}$$

$$K\_{sp}=[Pb\_{\left(aq\right)}^{2+}]^{1}×\left[I\_{\left(aq\right)}^{-}\right]^{2}$$

For lead(II) iodide, the solubility limit is determined to be 0.00135molL-1 at 25°C. For one mole of lead(II) iodide which enters solution you obtain one mole of lead(II) ions and two moles of iodide ions. This concentration is doubled for iodide, as per the equation above and produces the following calculation to find Ksp:

$$K\_{sp}=0.00135^{1}×(2×0.00135)^{2}$$

$$K\_{sp}=0.00135×0.00270^{2}$$

$$K\_{sp}=9.80×10^{-9}$$

To reverse this calculation from a known Ksp to determine a solubility using the same example of lead(II) iodide, from the data sheet where Ksp is 9.8x10-9. To make the mathematical display simpler we can let the solubility of lead(II) iodide be “*x*”:

$$PbI\_{2\left(s\right)} \leftrightarrow 1Pb\_{(aq)}^{2+}+2I\_{(aq)}^{-}$$

$$∴x \leftrightarrow 1x+2x$$

$$K\_{sp}=[Pb\_{(aq)}^{2+}]^{1}×[I\_{\left(aq\right)}^{-}]^{2}$$

$$where [Pb\_{\left(aq\right)}^{2+}]^{1}=x^{1} and [I\_{\left(aq\right)}^{-}]^{2}=2x^{2}$$

$$∴K\_{sp}=x^{1}×2x^{2}$$

$$K\_{sp}=4x^{3}$$

$$given K\_{sp}=9.80×10^{-9}$$

$$∴9.80×10^{-9}=4x^{3}$$

$$\frac{9.80×10^{-9}}{4}=x^{3}$$

$$\sqrt[3]{\frac{9.80×10^{-9}}{4}}=x$$

$$x= \sqrt[3]{2×10^{-9}}$$

$$∴x= 0.00135molL^{-1}$$

The solubility limit (maximum concentration) of lead(II) iodide and lead(II) ions is equal to *x* and the iodide ions is double this value:

[lead(II) iodide] $=x=0.00135M$

[lead(II) ions] $=x=0.00135M$

[iodide ions] $=2x=2×0.00135M=0.00270M$

More information on Ksp calculations can be found at [Chemistry LibreTexts – Solubility Equilibra](https://chem.libretexts.org/Bookshelves/General_Chemistry/Map%3A_Chemistry_-_The_Central_Science_%28Brown_et_al.%29/17%3A_Additional_Aspects_of_Aqueous_Equilibria/17.4%3A_Solubility_Equilibria).

### Calculating and applying Ksp in precipitation reactions

Ksp can be used to predict if any given reaction will produce a precipitate. It is important to remember that Ksp is temperature dependant and so, therefore, is solubility and the presence or absence of any precipitates at that temperature.

For this comparison we can again compare Ksp to Q. Where Q is greater than Ksp a precipitate will form, for Q is less than or equal to Ksp no precipitate will form. Q is calculated using the given solute concentrations to calculate the current conditions in the reaction. Comparing Q to Ksp shows if the solubility limit has been exceeded with the solute amount above the limit, the precipitate will be formed.

For example, will a precipitate form when 0.0010L of 0.0010M silver nitrate solution is added to 0.00010L of 0.0010M sodium chloride solution at 25oC?

Sodium nitrate is highly soluble (very high Ksp) and can be ignored as this would not contribute to any precipitate. Calculating the molarity of silver chloride produced using the concentration and volume data provided (calculating Q) where for silver chloride Ksp = 1.77×10-10:

$$Moles AgNO\_{3\left(aq\right)}=0.0010M×0.0010L=1.0×10^{-6} mol$$

$$Moles NaCl\_{\left(aq\right)}=0.00010M×0.0010L=1.0×10^{-7} mol$$

$$Reaction ratio is 1:1 so NaCl is a limiting reagent$$

$$Maximum moles AgCl\_{(s)} formed will be 1.0×10^{-7}moles$$

$$[AgCl\_{\left(s\right)}]=\frac{1.0×10^{-7}moles}{\left(0.0010L+0.00010L\right)}$$

$$\left[AgCl\_{\left(s\right)}\right]=9.1×10^{-5}$$

$$AgCl\_{\left(s\right)} \leftrightarrow 1Ag\_{(aq)}^{+}+1Cl\_{(aq)}^{-}$$

$$Q=[Ag\_{(aq)}^{+}]^{1}×[Cl\_{\left(aq\right)}^{-}]^{1}$$

$$Q=9.1×10^{-5}×9.1×10^{-5}$$

$$Q=8.3×10^{-9}$$

In this case, Q is greater than Ksp and so a precipitate will form under these conditions. More information on Ksp and solubility calculations can be found at [Chemistry LibreTexts – Criteria for Precipitation and its Completeness.](https://chem.libretexts.org/Bookshelves/General_Chemistry/Map%3A_General_Chemistry_%28Petrucci_et_al.%29/18%3A_Solubility_and_Complex-Ion_Equilibria/18.5%3A_Criteria_for_Precipitation_and_its_Completeness)

### Common ion effect

In cases where the solvent is another ionic solution rather than water, the solubility of another ionic substance is influenced by the common ion effect. This means the ions in solution will interact with the solute ions and if they are common it will reduce the overall solubility of the solute. Consider this example of lead(II) sulfate (Ksp = 2.53×10-8) in water and 2.0M sulfuric acid. Both the lead(II) sulfate and sulfuric acid have sulfate ions as common in solution, for lead(II) sulfate in water:

$$PbSO\_{4\left(s\right)} \leftrightarrow 1Pb\_{(aq)}^{2+}+1SO\_{4(aq)}^{2-}$$

$$∴x \leftrightarrow 1x+1x$$

$$K\_{sp}=[Pb\_{(aq)}^{2+}]^{1}×[SO\_{4\left(aq\right)}^{2-}]^{1}$$

$$where [Pb\_{\left(aq\right)}^{2+}]^{1}=x^{1} and [SO\_{4\left(aq\right)}^{2-}]^{1}=x^{1}$$

$$∴K\_{sp}=x^{1}×x^{1}$$

$$K\_{sp}=x^{2}$$

$$given K\_{sp}=2.53×10^{-8}$$

$$∴2.53×10^{-8}=x^{2}$$

$$\sqrt{2.53×10^{-8}}=x$$

$$∴x= 1.59×10^{-4}M$$

The solubility limit (maximum concentration) of lead(II) sulfate, lead(II) ions and iodide ions are equal in this dissociation, $1.59×10^{-4}M$. This concentration for sulfate ions is used in this next calculation, however, could be ignored as this concentration is so very minute compared to the sulfate ions contributed by the 2.0M sulfuric acid to make the calculation simpler. To calculate the concentration of the uncommon ion (in this case Pb2+), use the concentration of the common ion and the listed Ksp value for lead(II) sulfate:

$$PbSO\_{4\left(s\right)} \leftrightarrow 1Pb\_{(aq)}^{2+}+1SO\_{4 (aq)}^{ 2-}$$

$$H\_{2}SO\_{4\left(aq\right)} \leftrightarrow 2H\_{(aq)}^{+}+1SO\_{4 (aq)}^{ 2-}$$

$$\left[SO\_{4\left(aq\right)}^{2-}\right]=\left[SO\_{4\left(aq\right)}^{2-} from PbSO\_{4}\right]+\left[SO\_{4\left(aq\right)}^{2-} from H\_{2}SO\_{4\left(aq\right)}\right]$$

$$\left[SO\_{4 (aq)}^{ 2-}\right]=\left[1.59×10^{-4}M\right]+\left[2.0M\right]=2.000159M$$

$$K\_{sp}=[Pb\_{(aq)}^{2+}]^{1}×[SO\_{4\left(aq\right)}^{2-}]^{1}$$

$$[Pb\_{(aq)}^{2+}]^{1}=\frac{K\_{sp}}{[SO\_{4\left(aq\right)}^{2-}]^{1}}$$

$$[Pb\_{(aq)}^{2+}]^{1}=\frac{2.53×10^{-8}}{2.000159M} =1.2×10^{-8}M$$

The solubility limit (maximum concentration) of lead(II) sulfate, lead(II) ions and iodide ions are equal in this dissociation, $1.2×10^{-8}M$. From these calculations it is shown that the solubility of lead(II) sulfate decreases due to the common ion effect from $1.59×10^{-4}M$ in water to $1.2×10^{-8}M$ in a 2.0M solution of sulfuric acid. More information on the common ion effect and the impact on solubility calculations can be found at [Chemistry LibreTexts – Common-Ion Effect in Solubility Equilibria.](https://chem.libretexts.org/Bookshelves/General_Chemistry/Map%3A_General_Chemistry_%28Petrucci_et_al.%29/18%3A_Solubility_and_Complex-Ion_Equilibria/18.3%3A_Common-Ion_Effect_in_Solubility_Equilibria)

### Solubility and its use by ATSI peoples to remove Cycad toxins

The following paper is to be viewed, this paper is available free online through JSTOR:

[Beck, W. (1992). Aboriginal Preparation of Cycas Seeds in Australia. Economic Botany, 46(2), 133-147](https://www.jstor.org/stable/4255419?seq=1)

It describes and practically demonstrates the use of solubility equilibria to detoxify cycas seeds. The seeds are tested before and after treatment to measure the concentration of cycasin, the main toxic compound in cycas seeds. Cycasin can be visualised on [MolView](http://molview.org/internetExplorer) and chemical information obtained from [FoodB](https://foodb.ca/compounds/FDB018287) and [NIH PubChem](https://pubchem.ncbi.nlm.nih.gov/compound/Methylazoxymethanol-glucoside#section=MS-MS).

Cycasin has a solubility of 56.6gL-1 in water with many structural components that interact strongly with water. Once the cycasin is physically released from the plant material it can be extracted into solution and washed away. Several stages of preparation are described, and the results quantified with reliable analytical techniques.

### Student activity

Read through the paper and discuss with a peer, the class or the teacher. Construct a formal response to a question, such as: With reference to examples, including the use of solubility equilibria Aboriginal and Torres Strait Islander Peoples, describe how solubility relates to chemical equilibrium?

The following discussion questions can be considered in class as preparation:

* The ‘Zamia Staggers’ is a common symptom of cycas poisoning. What dosage limits are quoted for cycasin toxicity?
* How toxic are the cycas seeds tested before treatment? What limit for consumption would be possible?
* Discuss the evidence for, and effectiveness of, leeching as a method to reduce the toxicity of cycas seeds?
* Predict the reason for the importance of cycas seeds as a food source for Aboriginal People despite the risk posed by cycasin toxicity?