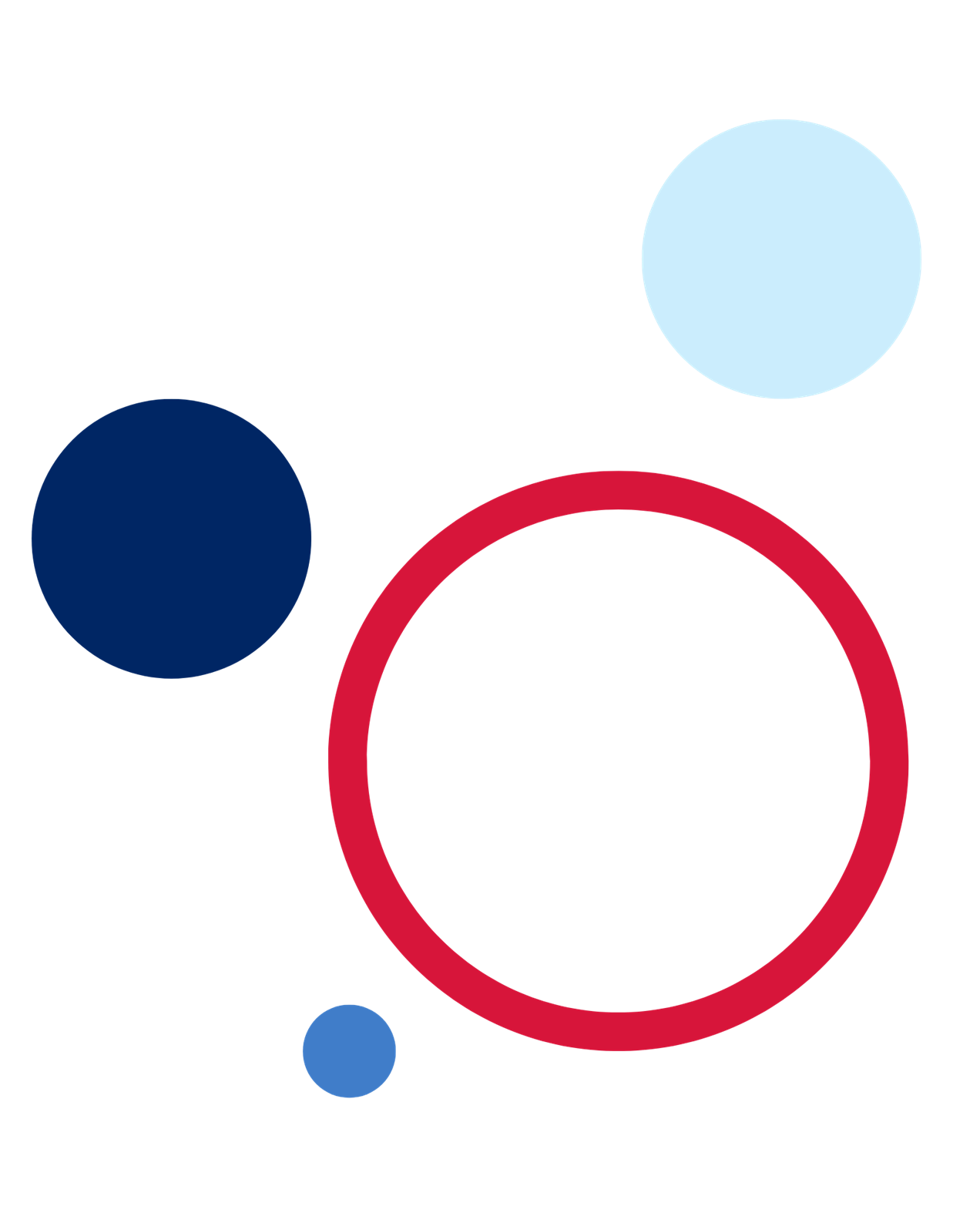
Chemistry Module 3: Reactive chemistry



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## Teaching the Year 11 Modules

The new Chemistry Stage 6 course was implemented in NSW schools in 2018. This syllabus incorporates new content and learning activities such as depth studies. The syllabus is designed around inquiry questions and formal assessment tasks emphasising Working Scientifically skills. The Year 11 course focuses on developing student knowledge and understanding of the relationships between the structure and the properties of atoms and molecules using macroscopic, microscopic, and symbolic representations. The Working Scientifically skills and processes are applied to predict outcomes by using physical, conceptual, and mathematical models and assessing the limitations of models. Through scientific investigations, students identify and measure quantities of chemicals, leading to a deep understanding of different classes of chemicals, including their structure, properties, and trends. The role of energy as a driver of chemical change and the relationship between energy and matter are explored in Year 11 as the basis for understanding chemical change in the Year 12 course.

As the syllabus is written as a series of inquiry questions and depth studies are a mandatory component of the course, pedagogies that promote student inquiry and deep learning should be employed in the chemistry classroom. New and challenging course content, and the change in pedagogical approach, were the catalysts for preparing these module guides for Stage 6. These guides are intended to assist teachers in delivering the chemistry syllabus effectively by outlining the big ideas, core concepts, strategies for teaching the modules, uncovering alternative conceptions, and providing strategies to address them and suggesting opportunities for extension. The guides support the teacher in facilitating the development of deep knowledge structures, such as the relationships between concepts. Teachers must note that the module guides are not a substitute for the syllabus but only support teachers to teach it. The module guides do not cover all aspects of the syllabus, as that was not within the project's scope.

The information contained in these documents is correct at the time of publication. While every effort has been made to eliminate errors, any errors or omissions identified after these documents are released will be corrected and released as resource updates. It is recommended that teachers access the [Curriculum website](https://education.nsw.gov.au/teaching-and-learning/curriculum/science/science-curriculum-resources-k-12) for the latest version of these documents.

## Course overview

The [Chemistry Stage 6 Syllabus](https://educationstandards.nsw.edu.au/wps/portal/nesa/11-12/stage-6-learning-areas/stage-6-science/chemistry-2017) explores the structure, composition and reactions of and between all elements, compounds and mixtures that exist in the Universe. The course develops an understanding of chemistry by applying Working Scientifically skills. It focuses on the exploration of models, the understanding of theories and laws, and the examination of the relationships between energy, matter, and change.

The chemistry course enables students to interpret the interconnection between the nature and practice of science and the knowledge of chemistry. The course addresses the discovery and synthesis of new compounds, the monitoring of elements and compounds in the environment, an understanding of industrial processes and their applications to life processes, our energy needs and uses, developing new materials, and sustainability issues.

The fundamentals developed in Year 11 include:

* Knowledge and understanding of the
* properties and structures of matter.
* relationship between observable properties and structures to trends in data and reactions.
* types and drivers of chemical reactions.
* role of energy in bonding within and between particles.
* using differing scales, measurements, and specialised representations to explain scientific phenomena and chemical interactions.
* skills in making hypotheses and designing valid and reliable investigations.
* skills in conducting investigations, solving problems, and explaining using cause and effect.
* constructing models, explaining phenomena using particle theory and discussing atomic 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.

## Module summary

See Appendix 2 for a table of key terms.

All chemical reactions involve the creation of new substances and associated energy transformations, commonly observable as changes in the temperature of the surroundings and/or the emission of light. These reactions are harnessed and controlled by chemists to produce substances that lead to the development of useful products.

Chemicals can react at different speeds and in many ways, yet they basically involve breaking and making chemical bonds. Students study how chemicals react, the changes in matter and energy that take place during these reactions, and how these chemical reactions and changes relate to the chemicals used in everyday life.

In this module, students focus on designing and conducting investigations to obtain and process data and evaluate the processes most appropriately in relation to chemical reactions. Students should be allowed to use all the Working Scientifically skills throughout the course.

## Big ideas

* The underlying chemical structure of substances is a key factor in determining their properties.
* Chemical systems are described using models based on particle theory. Models are revised as new evidence becomes available.
* Chemical bonding and reactions occur in predictable patterns according to their atomic and molecular structures.
* Energy changes can initiate chemical reactions. The underlying structure of the chemicals determines how they react to energy changes.
* Patterns can be determined by making observations, conducting fair tests, analysing data, and explaining trends in observable properties of substances and the Periodic Table.

## Relationship to other modules

There is some tension around the order for teaching parts of Modules 2 and 3. In Module 2, IQ2-1 (chemical reactions and stoichiometry) requires students to observe, measure and solve problems regarding mass changes in chemical reactions. However, Module 3, IQ3-1 (chemical reactions) is about chemical change, identifying products of various reactions and constructing balanced equations. Teachers may wish to teach IQ3-1 prior to or concurrently with Module 2 and must, therefore, carefully plan the learning sequence to ensure all parts of IQ3-1 are addressed. A deep knowledge of chemical reactions is necessary for students to be able to investigate Module 4, Drivers of Reactions.

* Students will link reactions of different metals with patterns in metal activity and atomic properties from Module 1. For example, in a Cu/Mg galvanic cell, why does the Mg oxidise preferentially to Cu?
* Students will link this learning to Module 1 for the chemical properties and bonding as drivers of chemical reactions and Module 4 for activation energy, noting that reactions do not all occur at the same rate and that a range of factors contribute to how fast a reaction will occur.

## Core concepts

* Students explore, in significant depth, a wide range of chemical reactions, including their products, reactants and stoichiometry.
* Students use definitions and patterns of chemical reaction types to describe and predict the reactants, products, and stoichiometry, for example, combustion reactions.
* Construction and operation of galvanic cells for REDOX reactions to occur.
* Students explore different reaction equation formats, including REDOX equations, half equations, and net ionic equations (by removing spectator ions).
* Change in reaction rate using collision theory (for reactant particles, the frequency, energy, and orientation of collisions). Factors can be deliberately controlled to manipulate a reaction rate.

## Opportunities for extending concepts

* Exploring chemical reactions that do not fall neatly into one category, such as hydrogen gas and oxygen gas, to produce water, is this oxidation or synthesis? Or is this combustion due to energy output?
* Contrasting synthetic and naturally occurring examples of the chemical reactions covered in this module.
* Detoxification of all foods through reactions during cooking, for example, denaturing proteins in raw foods and using lactase to remove for those who are lactose intolerant.
* Examples of reactions that could make foods toxic (instead of detoxification), including cooking bacon containing nitrates/nitrites, smoked or pickled foods and reheating foods.
* Exploring electrolytic cells by reversing a galvanic cell electron flow using an external power supply. For example, electroplating, restoring salvaged items, electrolytic (active) rust protection, electro bath paint and spray-painting electrical charge to encourage paint adhesion.
* Application of REDOX reactions in different types of cells, for example, hydrogen fuel cells, as a replacement for the internal combustion engine.
* Quantitative rates of reactions and rate constants. Making stoichiometric measurements of reaction rate.

## EAL/D strategies

One of the most significant challenges in Module 1 is the large amount of new terminology students are exposed to. Students must be explicitly taught the skills necessary to develop a glossary that can be consistently added to in Module 1 and subsequent modules.

* Explore different methods of storing glossary terms with students, such as using physical flashcards or one of the many digital flashcard apps available ([Brainscape](https://www.brainscape.com/), [Cram](https://www.cram.com/), [Quizlet](https://quizlet.com/latest),.
* Model for students how to define terms using their own words.
* Encourage students to use analogies and metaphors when appropriate and include diagrams/images.

Videos such as [Chemical reactions introduction | Chemistry of life | Biology | Khan Academy (9:12)](https://www.youtube.com/watch?v=TStjgUmL1RQ) can define this terminology, and regularly discussing terminology will reinforce learning. Sentence frames can assist students in constructing sentences independently. Formative assessments and short responses, designed to allow students to practice using technical language, link ideas and construct explanations and cause and effect relationships, should be frequent with feedback regarding student use of technical language and complex sentences. Teaching the root of words can help students decode unfamiliar words.

## Misconceptions

* Students may perceive chemical reactions as occurring instantly and directly converting reactants to products. Concepts of intermediate products or slower reaction rates are unfamiliar to these students. By undertaking practical investigations of clock reactions, for example, the [iodine clock](https://edu.rsc.org/experiments/iodine-clock-reaction-demonstration-method/744.article) and [traffic lights](https://edu.rsc.org/exhibition-chemistry/traffic-lights/2020080.article), students can observe these slower reactions, many of which involve intermediate products.
* Students may not associate making and breaking bonds with electron transfers. Bonds and electrons are often considered separate concepts. Creating half equations to show electron changes in chemical reactions, including changes to oxidation numbers, will show electron movement as a key step in the bonding process.
* Students may experience confusion with equations showing half ratios in their balancing ratios, such as the complete combustion of ethane:

Until now, balancing ratios have typically appeared as integer values only. The typical use of half values in combustion reactions is to establish the standard form of the equation using one mole of fuel to link the balanced equation with the enthalpy of the combustion value. This unique use of non-integer balancing values helps explain the differences between complete and incomplete reactions in terms of enthalpy and mole ratios. However, this is not typically used elsewhere.

* Students may not recognise the differences between complete and incomplete combustion. Their previous experiences have likely not included a separation between these two types of reactions and the differences in the energy and reaction products. This can be addressed by using examples of both complete and incomplete combustion reactions for calorimetry experiments, for example, the two flame settings on a Bunsen burner. Students will clearly observe a slower heating rate and the production of soot (carbon deposits) with incomplete combustion reactions.
* When students consider precipitation reactions, they may use the “swapping partners” analogy, where students believe all ions partner up in solution to produce insoluble (precipitates) or soluble compounds. Changing the way precipitation equations are written to show all soluble species as ions will assist students in recognising the importance of the solvent. We can only bring the soluble ions together as a compound by removing the solvent. In the net ionic equations, these ions are spectators and redacted. For example, the full ionic equation for the precipitation of silver chloride:

This can be simplified using the net ionic equation where the spectator ions (sodium and nitrate) are removed:

* Some reactions of metals are too slow to observe, and students may not recognise that a reaction is occurring. Magnesium reacts slowly with water, and aluminium has a passivating layer which retards normal reactions. This can be addressed by using practical examples to challenge student preconceptions and include ways to measure slow reactions, for example, pH indicators to measure the pH rise as magnesium reacts with water and [ferroxyl indicator](https://edu.rsc.org/exhibition-chemistry/nailing-corrosion-demonstrations/2000054.article) to show REDOX reactions in metal activity practicals visually.
* Most students have used sandpaper with metal samples before use. The misconception is that the metal needs to be ‘cleaned’ to be reactive instead of recognising the presence of corrosion coatings. This can be addressed by being very clear with students about the planning and conducting of metal activity series and galvanic cell practicals to obtain valid and reliable results.

Conceptual difficulties

* Some indicators of chemical change are not as clear as students may expect, for example, precipitation reactions forming suspensions rather than sediments. Students may believe a reaction is not occurring due to an overly rigid definition of indicators for chemical change.
* The significant vocabulary requirements to describe the components and processes of a REDOX equation – anode/cathode, oxidation/reduction, positive/negative, oxidant/ reductant, oxidising agent/reducing agent. The diagram on page 33 can assist students in navigating these terms and their relationship to each other.
* Students who study chemistry and physics will experience a conflict in the terminology of cathode/anode, where this module addresses galvanic cells and physics Module 8 applies the same terminology to electrolytic processes. Due to the inverse nature of each process, the terminology is also inverted.
* Students need to recognise that an inert metal electrode in galvanic cells does not take part in chemical reaction, for example, the platinum electrode in the standard hydrogen electrode.
* Standard reduction potentials are measured under standard conditions. Students may question experimental results which differ from the datasheet. It is important to reinforce the nature of standard conditions for reference data measurement, which can also allow students to explore the impact of variables on these measurements.
* Students may find it hard to reconcile the meaning of a negative standard potential. By reviewing the reactions with negative values, students can develop an understanding that the reduction is increasingly less likely to occur, having a preference for oxidation and the reverse of the reaction shown.
* When considering reaction rates, some chemical reactions have such a high or low rate that they are difficult to interpret. Students need to experience chemical reactions with reasonable time scales to measure them and to fit the length of time they have in class to experience them.
* Students may confuse catalysts as reactants or products and attempt to include them in a chemical equation, creating difficulty balancing them. Reminding students that any catalysts are represented as a condition above the reaction arrow and do not need to be considered when balancing the equation.

## Suggested teaching strategies

### IQ3-1 Chemical reactions

#### Chemical change

Writing chemical equations in documents can be difficult. [Writing Chemical Equations on the Web from Periodni](https://www.periodni.com/chemical_equations_on_web.php) is an online tool that can provide a useful shortcut to copy and paste chemical equations into other applications and formats.

Chemical changes cause a change in the chemical bonding between substances, a chemical reaction. These changes can only be reversed by additional chemical reactions to reverse the change in chemical bonding.

Indicators of chemical changes can be summarised as:

|  |  |  |
| --- | --- | --- |
| Sign of chemical change | Examples of chemical change | Everyday example(s) |
| Temperature rises (exothermic reaction) | Magnesium and hydrochloric acid: | Cellular respiration and combustion |
| Temperature falls (endothermic reaction) | Aluminium chloride and barium hydroxide: | Photosynthesis and instant icepacks |
| Appearance of a new product | Sodium chloride and silver nitrate produce silver chloride precipitate: | Limescale from hard water and rust on metal objects |
| Disappearance of a reactant | Calcium carbonate and hydrochloric acid: | Formation of limestone caves and combustion |
| Permanent colour change | Carbonisation of white sucrose using sulfuric acid produces black carbon and water (as steam): | Maillard reaction when cooking meat |
| Evolution of a gas product (bubbling or effervescence) | Calcium and water produce hydrogen and calcium hydroxide: | Sherbet and fermentation in rising bread |
| Production of a new odour | Release of hydrogen sulfide gas from the reaction of iron(II) sulfide with hydrochloric acid: | Decomposition of foods and flatulence |

Making and separating mixtures are physical changes. Dissolving a soluble substance in a solvent is not a chemical change. The original solute is still detectable and can be recovered using physical processes instead of chemical reactions.

#### Balancing equations

Many techniques can be used to balance chemical equations. By showing students a range of methods, they are better placed to select the most appropriate method for their needs and develop proficiency in this core chemistry skill. Many online tools are available. For example, [Balancing Chemical Equations from PhET](https://phet.colorado.edu/en/simulation/balancing-chemical-equations) or [Chemical Equations from Simpop](https://simpop.org/reactions/reactions.htm) can assist students in practising using a range of examples.

Conversion of names to chemical formulas

It will be impossible to balance the chemical equation correctly without an accurate molecular formula for all the components.

Converting the word equation into a chemical equation using appropriate molecular formulae is always required. Word equations do not allow stoichiometric calculations of reactants and products. This step is also a common feature of marking schemes for students to produce a correctly balanced chemical equation. This process is covered by content descriptors in Module 1 and is described in the [Module 1 guide [DOCX 1.4 MB]](https://education.nsw.gov.au/content/dam/main-education/teaching-and-learning/curriculum/key-learning-areas/science/s-6/chemistry/Chemistry-module-1-guide.docx).

Coefficients

When a correct molecular formula is present in a reaction equation without any integer at the front indicates a value of one, the value of one is always omitted:

One water molecule contains two hydrogen atoms and one oxygen atom.

Integers greater than 1 are placed at the front to indicate a multiple of all atoms in the formula. This value is then used to multiply all atoms in the molecular formula, a strategy to assist students is to allow them to add brackets (not parentheses, as these could be confused with features of some molecular formulas) around the entire formula. This will be more familiar to students as a related strategy with the use of grouping symbols in mathematics where the integer is multiplied through all components of the brackets:

Two water molecules containing (in total) four hydrogen atoms and two oxygen atoms.

This can become more complex with polyatomic molecular structures. Careful construction and a clear layout of the mathematical notation are vital to ensure clarity to students, for example, calcium phosphate:

Three calcium phosphate formula units containing (in total) 9 calcium atoms, 6 phosphorus atoms and 24 oxygen atoms.

Coefficients can be fractions only in combustion reactions where it is standard practice to ensure the fuel is shown as one mole in the reaction equation.

Balancing by inspection

Simple equations can be balanced by direct inspection. The Haber process is the production of ammonia from nitrogen and hydrogen, as a word equation:

Converting to the word equation to an unbalanced chemical equation:

Work from the largest molecule by the number of atoms in the products.

Balance the first atom nitrogen (N), gives a coefficient of 2 for ammonia (NH3) required to meet the needs of the nitrogen (N2)in the reactants:

Moving to the next atom, hydrogen (H), as H3 multiplied by the existing coefficient of 2, becomes 6x hydrogen atoms in the products. Therefore, a coefficient of 3 must be placed in front of the hydrogen gas (H2) reactant to balance hydrogen for conservation of mass:

This chemical equation is now balanced.

If the reaction equation includes oxygen atoms, it may be easier to leave it till last when balancing, as it is frequently part of multiple formulae in the equation. By balancing the other elements and polyatomic ions, oxygen will commonly self-balance.

##### Tabulation

More complex equations can be tabulated or tallied with the atom counts in the reactants and products tracked throughout the changes made to the coefficients. Although this method may be slow to begin with, by starting with simple reactions, students can develop proficiency to undertake this operation mentally or with minimal shorthand notes.

Many pathways can be taken to balance each element in the equation. Each has its benefits and limitations. A suggested order could be:

1. Metals
2. Polyatomic ions (which are retained unchanged in the products)
3. Non-metals (except hydrogen and oxygen)
4. Hydrogen
5. Oxygen (which will hopefully self-balance from steps 1-4)

For example, when heated, solid aluminium hydroxide decomposes into solid aluminium oxide and water vapour:

Converting this word equation to an unbalanced chemical equation:

Create a table to tally the number of each element on both sides of the equation. For this example, the order suggested above has been used from top to bottom (1-5):

|  |  |  |  |
| --- | --- | --- | --- |
| Component | Reactants | Products | Balanced |
| Al | 1 | 2 | Not balanced |
| H | 3 | 2 | Not balanced |
| O | 3 | 4 | Not balanced |

Starting with aluminium (metals), we are required to double the reactants side to bring equal with products and adjust these amounts in the table:

|  |  |  |  |
| --- | --- | --- | --- |
| Component | Reactants | Products | Balanced |
| Al | 2 | 2 | Balanced |
| H | 6 | 2 | Not balanced |
| O | 6 | 4 | Not balanced |

This equation does not include any polyatomic ions retained through to the products or other non-metals (hydrogen is dealt with separately). Next, hydrogen will be balanced. We have to bring the products up to 6. To do this, we can multiply water (containing two hydrogens each) by 3:

|  |  |  |  |
| --- | --- | --- | --- |
| Component | Reactants | Products | Balanced |
| Al | 2 | 2 | Balanced |
| H | 6 | 6 | Balanced |
| O | 6 | 6 | Balanced |

By taking these steps, oxygen has also self-balanced through its inclusion in various other molecular formulas. Now that all elements are balanced, the equation is balanced.

##### Polyatomic ions

Polyatomic ions are cation or anion groups that consist of a combination of elements. In some reactions, these polyatomic ions survive into products unchanged. In other reactions, they are decomposed.

Where a polyatomic ion is retained in the products, it can be treated as a single species to balance the equation.

For example, when an aluminium sulfate solution reacts with a calcium hydroxide solution, it produces an aluminium hydroxide solution and a calcium sulfate solid precipitate:

Converting this word equation to an unbalanced chemical equation and creating a tally table for each component of the chemical equation:

Create a table to tally the number of each element on both sides of the equation. For this example, the suggested order has been used from top to bottom:

|  |  |  |  |
| --- | --- | --- | --- |
| Component | Reactants | Products | Balanced |
| Ca | 1 | 1 | Balanced |
| Al | 2 | 1 | Not balanced |
| SO4 | 3 | 1 | Not balanced |
| OH | 2 | 3 | Not balanced |

Balancing top to bottom, calcium is already balanced (for now), and the aluminium in the products needs to double. Placing the coefficient of 2 in front of the product containing aluminium will balance this component, noticing this also doubles the hydroxide in the products:

|  |  |  |  |
| --- | --- | --- | --- |
| Component | Reactants | Products | Balanced |
| Ca | 1 | 1 | Balanced |
| Al | 2 | 2 | Balanced |
| SO4 | 3 | 1 | Not balanced |
| OH | 2 | 6 | Not balanced |

Moving next to the unbalanced sulfate, placing a 3 on the sulfate product brings this component to being balanced, noticing this also triples the calcium in the products:

|  |  |  |  |
| --- | --- | --- | --- |
| Component | Reactants | Products | Balanced |
| Ca | 1 | 3 | Not balanced |
| Al | 2 | 2 | Balanced |
| SO4 | 3 | 3 | Balanced |
| OH | 2 | 6 | Not balanced |

As the calcium is now unbalanced, return to the top of the table, and place the coefficient of 3 on the calcium of the reactants, noticing this also triples the hydroxide in the reactants:

|  |  |  |  |
| --- | --- | --- | --- |
| Component | Reactants | Products | Balanced |
| Ca | 3 | 3 | Balanced |
| Al | 2 | 2 | Balanced |
| SO4 | 3 | 1 | Balanced |
| OH | 6 | 6 | Balanced |

This chemical equation is now balanced with all reaction components in the same amounts across the reactants and the products.

Where a polyatomic is decomposed in the reaction, it must be balanced as separate elements. Some shortcuts are possible.

For example, when sodium carbonate solid reacts with a hydrochloric acid solution, it produces a sodium chloride solution, carbon dioxide gas and water:

Converting this word equation to an unbalanced chemical equation and creating a tally table for each component of the chemical equation:

|  |  |  |  |
| --- | --- | --- | --- |
| Component | Reactants | Products | Balanced |
| Na | 2 | 1 | Not balanced |
| C | 1 | 1 | Balanced |
| Cl | 1 | 1 | Balanced |
| H | 1 | 2 | Not balanced |
| O | 3 | 3 | Balanced |

Balancing top to bottom, the sodium in the products needs to double. Placing the coefficient of 2 in front of the product containing sodium will balance this component, noticing this also doubles the chloride in the products:

|  |  |  |  |
| --- | --- | --- | --- |
| Component | Reactants | Products | Balanced |
| Na | 2 | 2 | Balanced |
| C | 1 | 1 | Balanced |
| Cl | 1 | 2 | Not balanced |
| H | 1 | 2 | Not balanced |
| O | 3 | 3 | Balanced |

Carbon is already balanced (for now). Looking down the table at chloride, the reactants need to double by placing the coefficient of 2 for the reactant chloride, noticing this also doubles the hydrogen in the reactants:

|  |  |  |  |
| --- | --- | --- | --- |
| Component | Reactants | Products | Balanced |
| Na | 2 | 2 | Balanced |
| C | 1 | 1 | Balanced |
| Cl | 2 | 2 | Balanced |
| H | 2 | 2 | Balanced |
| O | 3 | 3 | Balanced |

This chemical equation is now balanced with all reaction components in the same amounts across the reactants and the products.

For simple acid/carbonate reactions, a generalisation for students to use as a shortcut is to notice that the coefficients for carbon dioxide and water are equal to the number of carbonate ions in the reactants.

Alternative methods using [algebraic techniques](https://www.chembuddy.com/?left=balancing-stoichiometry&right=algebraic-method) can be explored for those students who are more mathematically capable:

* [How to balance chemical equations algebraically (12:25)](https://www.youtube.com/watch?v=kjNcL92nocM)
* [Chemistry: Balancing Chemical Equations (algebraic method) (14:02)](https://www.youtube.com/watch?v=tUBzpsvMeWg)
* [01 – Introduction to the Algebraic Method for Balancing Chemical Equations (3:55)](https://www.youtube.com/watch?v=eS9CmMFbMXg&t)
* [02 – Algebraic Method for Balancing Chemical Equations (Example #1) (3:24)](https://www.youtube.com/watch?v=6kDc6Jb44D8)
* [03 – Algebraic Method for Balancing Chemical Equations (Example #2) (8:11)](https://www.youtube.com/watch?v=Q_wSoBP3jk4).

#### Types of chemical reactions

Classifying chemical reactions is a useful strategy to show the common features of reaction types. This allows students to quickly deduce information from the simple inspection of a chemical reaction equation. The list of reaction types included is not exhaustive, and many other types exist for possible depth studies.

Many chemical reactions will not neatly exist in one category and exhibit the features of multiple types of reactions described here.

##### Synthesis

In a synthesis reaction, two or more reactants combine to form an aggregate product. In general terms, it can be described as:

For example, the production of ammonia from nitrogen gas and hydrogen gas (the Haber process):

Practically, students can experience synthesis reactions by burning magnesium ribbon in a crucible to form white magnesium oxide powder:

Alternatively, the ‘pop test’ for Hydrogen gas to produce water (this synthesis reaction is also a combustion reaction):

##### Decomposition

A decomposition reaction is the inverse of a synthesis reaction. This is where a reactant breaks down into two or more products. In general terms, it can be described as:

For example, the electrolysis of water to form hydrogen gas and oxygen gas:

Practically, students can experience decomposition reactions by heating green copper carbonate in a test tube using a Bunsen burner to produce black copper oxide and carbon dioxide gas (which can be collected using a side-arm test tube and verified using a limewater trap):

The decomposition of hydrogen peroxide ([many catalysts are possible](https://edu.rsc.org/resources/hydrogen-peroxide-decomposition-using-different-catalysts/831.article) and could form an investigation) to form water and oxygen gas as a teacher demonstration and/or selected online videos, for example, [Genie in a bottle (3:58)](https://www.youtube.com/watch?v=5q5bzHckSIM) or [elephants' toothpaste (4:05)](https://www.youtube.com/watch?v=p1eG2y2mn54) (taking this to the extreme is the [world record example (2:06)](https://www.youtube.com/watch?v=IyDt4qXTjBM)):

##### Combustion

Combustion is the reaction between a substance (fuel) and an oxidant (usually oxygen gas), releasing energy (exothermic) and producing an oxidised product of the fuel. There is an overlap with the concept of corrosion (oxidation), an example of a naturally occurring slow reaction of substances with oxygen gas. Although the most common example is the combustion of hydrocarbons, alternative examples are important for students to recognise as combustion, for example, the combustion of sulfur with oxygen gas to form sulfur dioxide (caution with the sulfur dioxide gas produced):

Balancing oxygen using fractional coefficients is commonly used in combustion reactions to standardise the equation as one fuel molecule undergoing complete combustion.

In a more traditional sense, the combustion of hydrocarbon fuels is the stereotypical combustion reaction. For example, octane (a component of petrol) is combusted in the engine using oxygen from the air to produce carbon dioxide gas and water vapour:

A range of combustion reactions can be practically explored by students or through teacher demonstrations. For example, ethanol undergoes combustion with oxygen to produce carbon dioxide and water:

Allowing students to explore combustion reactions using a range of fuels will allow them to stumble across and attempt to explain the concepts of incomplete combustion through observation. For example, the comparison to the incomplete combustion of ethanol (under reduced oxygen conditions) to produce carbon soot and water:

By providing opportunities to include a basic form of calorimetry to measure the combustion energy, this activity will provide a useful introduction to the work, which will be completed through Module 4 and Module 7.

##### Precipitation

Precipitation reactions involve the combination of two soluble substances, which then form at least one insoluble substance through the rearrangement of ions. Generally, it is shown as:

For example, where one precipitate forms:

Precipitation reactions can also produce two insoluble products, for example (magnesium hydroxide is only slightly soluble):

Students can practically explore various reactions to create many cation and anion combinations. Through these observations, they can deduce some rules for cation/anion combinations that are insoluble and compare these to a solubility reference table. This learning activity will continue to be useful to students as they progress through Module 8.

The [data sheet [PDF 880 KB]](https://educationstandards.nsw.edu.au/wps/wcm/connect/98664936-221f-4c49-88e1-d002ec69285c/chemistry-formulae-sheet-data-sheet-periodic-table-hsc-exams-2019.pdf?MOD=AJPERES&CVID=) contains a list of solubility constants, which can also be used as a reference table for a list of insoluble substances.

Students can practically explore these precipitation reactions and combine them with a quantitative gravimetric analysis. For example:

* The precipitation and quantification of sulfate in lawn fertiliser is a typical example of this type of analysis. This activity is described in Appendix 1 of the [Module 8 guide [DOCX 448 KB]](https://education.nsw.gov.au/content/dam/main-education/teaching-and-learning/curriculum/key-learning-areas/science/s-6/chemistry/CHEM_Module_8_Guide.docx).
* The colourful [golden rain (11:06)](https://www.youtube.com/watch?v=4xSUvSGywXo) precipitation of lead (II) iodide is visually pleasing but results in significant amounts of lead waste. This activity would be best undertaken as a teacher demonstration only.

##### Acid/base reactions

Acids and bases undergo neutralisation reactions to produce salt and water. Generally, using a hydroxide as the base:

An example of this is the reaction of Hydrochloric acid with Sodium Hydroxide:

Using pH indicators and/or pH probes will give students a visual indicator of the progression of these chemical reactions, including a method to quantify the relationship between acids and bases based on the pH scale from prior knowledge in stage 5. This type of activity will provide a useful introduction to the work completed in Module 6.

##### Acid/carbonate reactions

Acids also react in characteristic ways with carbonates. This is another type of acid/base reaction and is distinguished by its production of carbon dioxide gas:

An example of this is the reaction of Hydrochloric acid with Sodium Hydrogen Carbonate:

Other practical opportunities to explore with students could include using these reactions in real-world examples, such as making sherbet, baking cakes, and modelling the production of cave systems by reacting limestone chips with acids.

#### Detoxification of foods

This content descriptor relates the chemical processes in this module to the methods used by Aboriginal and Torres Strait Islander Peoples to detoxify poisonous food items. By integrating a range of contemporary methods derived from traditional knowledge, students can connect traditional knowledge to first-hand experience with modern practices and explain them using chemical reaction types and terminology.

Many food products consumed by humans contain chemical compounds or [potentially pathogenic organisms](https://www.foodstandards.gov.au/publications/pages/compendium-of-microbiological-criteria-for-food.aspx) which can cause harm to the consumer. Processes to sterilise, cook, pack and store foods are used to reduce the instances of food poisoning that can result from these compounds or pathogens. Many examples of the proper preparation of foods are presented in the [Safe Food Australia](https://www.foodstandards.gov.au/publications/pages/safefoodaustralia3rd16.aspx) guide. For example:

* Raw chicken can contain [Salmonella and Campylobacter](https://www.foodstandards.gov.au/science/surveillance/pages/baselinesurveyonthep4998.aspx), which can cause gastroenteritis. These bacteria are destroyed with proper cooking temperatures.
* [Uncooked rhubarb](https://www.compoundchem.com/2015/04/16/rhubarb/) contains oxalic acid and oxalates, which can cause kidney damage. These compounds are decomposed with proper cooking temperatures.
* Many [types of fungi](https://fungimap.org.au/about-fungi/edible-poisonous-fungi/) contain a wide range of poisonous compounds. Some species contain compounds resistant to all cooking processes and never lose their toxicity, so they cannot be consumed without harm.

Aboriginal people had a range of foods that required some form of treatment before consumption to reduce the instances of poisoning. The most well-known example, which will also be addressed in Module 5 (using solubility equilibria), is the detoxification of Cycas seeds. The seeds are harvested and treated using a range of methods described in [Aboriginal Preparation of Cycas Seeds in Australia](http://www.jstor.org/stable/4255419) to reduce the Macrozamin and Cycasin content of the plant material. Further reading on these compounds can be found in [The Role of Macrozamin and Cycasin in Cycads (Cycadales) as Antiherbivore Defenses](https://www.jstor.org/stable/3557555).

Figure 1 – mature cycad fruits



"[Cycad](https://www.flickr.com/photos/22616984@N07/2423489087)" by [Ian Sutton](https://www.flickr.com/photos/22616984@N07/) is licensed under [CC BY-NC-SA 2.0](https://creativecommons.org/licenses/by-nc-sa/2.0/).

### IQ3-2 Predicting reactions of metals

#### Reactions of metals

Metals are the dominant classification of elements on the Periodic Table. The chemical reactions of each metal to the reactants and conditions described in this content descriptor depend on a range of factors described later in this inquiry question.

##### Water

Active metals, particularly the group 1 (alkali metals) and group 2 (alkaline earth metals) metals of the Periodic Table, will vigorously react with water to produce their corresponding hydroxide and hydrogen gas. In general terms:

For example, sodium metal reacts with water to produce sodium hydroxide and hydrogen gas:

In the case of group 1 metals, this reaction can produce enough energy to ignite the hydrogen gas and [cause an explosive reaction (5:42)](https://www.youtube.com/watch?v=mzEOL-N2cbw). For this reason, the only permitted demonstration of this type of reaction uses sodium metal. Extreme care must be taken for this teacher demonstration to demonstrate this reaction to students safely. Please refer to [sodium metal in CSIS](https://ecmjsp.education.nsw.gov.au/ecmjsp/chemicals/?mode=viewchemares&chemalpha=S&chemid=1524#skipToContent) for guidance. The other group 1 metals are too reactive to store and use in an educational setting safely and are not permitted for use in schools.

For student investigations, the same reaction can be observed with [calcium metal](https://ecmjsp.education.nsw.gov.au/ecmjsp/chemicals/?mode=viewchemares&chemalpha=C&chemid=397#skipToContent) as it will not typically result in the same explosive reaction as sodium metal. Calcium metal has a higher density than water and will sink rather than float like sodium. Some considerations when planning this activity include:

* [Collecting the hydrogen gas](https://edu.rsc.org/experiments/generating-collecting-and-testing-gases/693.article) generated in a water-filled, inverted measuring cylinder to quantify the gas volume to support stoichiometric calculations with this reaction.
* The addition of phenolphthalein to observe the increase in pH indicated by the vibrant pink colour produced by hydroxide ions.
* Investigating the influence of water temperature on the rate of reaction, measured by the rate of hydrogen gas production. Limiting the temperature of the water to a maximum of .

Less active metals will react very slowly or not at all with water but may react with steam. In this case, the metal salt produced will be an oxide instead of a hydroxide, as found in the more active metal examples:

For example, the reaction of Magnesium metal with steam:

##### Dilute acid

Most metals will react with dilute acids, particularly the active metals, to produce their corresponding salt and hydrogen gas. In general terms:

For example, magnesium metal reacts with hydrochloric acid to produce magnesium chloride and hydrogen gas:

Due to the vigorous nature of these reactions, it is not suitable to conduct this reaction with more active metals, for example, sodium.

For student investigations, these reactions can be observed with a range of metals and acids to compare the relative reaction rates, as measured by the time taken to consume the equimolar quantity of the metal or the rate of hydrogen gas production.

Students will also become familiar with the relative strengths of acids and reaction ratios for monoprotic, diprotic and triprotic acids by investigating the role of different acids. This will assist students as they continue through related concepts in Module 6.

It is not suitable to conduct this reaction with concentrated oxidising acids due to the production of nitrogen dioxide when using nitric acid and sulfur dioxide when using sulfuric acid. Other non-oxidising acids will produce highly exothermic reactions that may result in the detonation of the hydrogen gas produced.

##### Oxygen

Most metals will react slowly with oxygen. The more active metals will react quickly to produce their corresponding oxide. In general terms:

For example, aluminium metal reacts with oxygen to produce aluminium oxide:

As a teacher demonstration, a freshly cut sliver of sodium metal will appear lustrous and become dull as the surface reacts quickly with atmospheric oxygen. Students will be able to investigate the reaction of [magnesium metal](https://ecmjsp.education.nsw.gov.au/ecmjsp/chemicals/?mode=viewchemares&chemalpha=M&chemid=1015#skipToContent) (ignited by a blue Bunsen flame) and the production of white magnesium oxide powder. This reaction can be [completed in a crucible](https://edu.rsc.org/experiments/the-change-in-mass-when-magnesium-burns/718.article) to carefully capture and quantify this reaction by mass change to validate the stoichiometry.

Some metals will react so slowly with oxygen that students cannot observe. This slow rate of reaction can be related to the corrosion resistance of the metals in the environment, a useful property for the intended end-use of the metal.

##### Other metal ions in solution

All metals, except for noble metals, for example, gold and platinum, will react with a range of metal ion solutions according to their relative strength as a reductant (ability to cause reduction of the metal ion species in solution). Generally, more active metals (Metal A) will displace a less active metal ion (Metal B ion) in solution through an exchange of electrons, resulting in the oxidation of Metal A and reduction of Metal B:

For example, magnesium metal will displace copper ions from the solution, resulting in the production of magnesium ions and copper metal:

Various metals and metal ion combinations will undergo a chemical reaction. Students could explore all combinations using available resources to collect data on relative metal reactions. Due to the high costs of silver nitrate, a suitable teacher demonstration could be the [microscopy analysis](https://edu.rsc.org/exhibition-chemistry/displacement-reaction-of-silver-nitrate-and-copper-metal/2020046.article) of silver crystal formation and the [silver Christmas tree (1:14)](https://www.youtube.com/watch?v=yO9sl60XAZo) activity.

By using the standard potentials [datasheet [PDF 880 KB]](https://educationstandards.nsw.edu.au/wps/wcm/connect/98664936-221f-4c49-88e1-d002ec69285c/chemistry-formulae-sheet-data-sheet-periodic-table-hsc-exams-2019.pdf?MOD=AJPERES&CVID=) in combination with observations of metal/metal ion reactions, students could determine the relationship (and even make predictions) between the order of this list and the metals' capacity to displace metal ions from solution:

Metals (as solids) higher on the standard potentials list (increasingly negative potentials) can displace metal ions lower on the list (increasingly positive potentials).

#### Metal activity series

Using the reactions described above, a combination of the data can be used to develop a relative scale of reactivity. Once completed, this relative scale can be compared to the standard potentials list on the [datasheet [PDF 880 KB]](https://educationstandards.nsw.edu.au/wps/wcm/connect/98664936-221f-4c49-88e1-d002ec69285c/chemistry-formulae-sheet-data-sheet-periodic-table-hsc-exams-2019.pdf?MOD=AJPERES&CVID=). [Online simulations](http://chemcollective.org/oxredux) are available to support the range of laboratory investigations undertaken.

##### Ionisation energy

The ionisation energy is the energy required to remove the most loosely held electron from one mole of gaseous atoms or ions to produce one mole of gaseous ions of the element.

There are several successive ionisation energies possible:

* First ionisation energy – removing one mole of electrons from one mole of the gaseous atom to produce a single positive charged cation.
* Second ionisation energy – removing one mole of electrons from one mole of the gaseous cations to produce a double positive charged ion.
* Nth ionisation energy – removing N moles of electrons from N moles of the gaseous cations to produce a Nth positive charged ion.

Metals with a lower first ionisation energy will exhibit a higher reactivity due to the reduced energy requirement for an electron transfer to occur. This produces more vigorous reactions with a wider range of chemical environments.

This concept links students back to Module 1, where they reviewed the periodic trends with ionisation energy. Students can draw upon this learning to relate the position of the metal on the Periodic Table to its observed trends in reactivity and to evaluate any contradictory evidence they may uncover during this process.

##### Atomic radius

Atomic radius is frequently defined as the distance between the nucleus and the outermost electron. Although this simplified definition is suited for general use, it can be problematic when describing ions where the atomic radius decreases for cations and increases for anions of the same element. A more suitable definition would be:

Atomic radius is half the distance between the nuclei of two adjacent identical atoms.

The atomic radius can be related to ionisation energy by the relative strength of attraction between the nucleus and the outermost valence electrons of the metal. These outermost electrons will first interact with the chemical environment into which the metal is placed.

For metals, the larger the atomic radius, the lower the first ionisation energy. The valence electrons are less tightly bound to the nucleus and are, therefore, more easily removed by interactions with other species within their chemical environment. The larger the atomic radius, the more reactive the metal will be.

This concept links students back to Module 1, where they reviewed the periodic trends with atomic radius. Students can draw upon this learning to relate the position of the metal on the Periodic Table to its observed trends in reactivity and to evaluate any contradictory evidence they may uncover during this process.

##### Electronegativity

Electronegativity can be easily confused with [electron affinity](https://chem.libretexts.org/Bookshelves/Physical_and_Theoretical_Chemistry_Textbook_Maps/Supplemental_Modules_(Physical_and_Theoretical_Chemistry)/Physical_Properties_of_Matter/Atomic_and_Molecular_Properties/Electron_Affinity) through the use of simplified definitions, for example, the strength of attraction between a nucleus and an electron. A more suitable definition would be:

Electronegativity is the ability of single atoms to attract electrons from outside, measured in ‘Pauling units’. This is a relative scale that allows atoms to be compared on their ability to attract electrons in the chemical bonds formed between similar or dissimilar atoms.

Metals typically have a lower electronegativity when compared to non-metals, and yet there are several highly reactive metals. Chemical reactions involve electron exchange, either acceptance or donation. Typically, metals participate in chemical reactions by donating electrons, and non-metals accept electrons. The concept of electronegativity as an electron-attracting force needs to be contrasted with an opposite electron-repelling force (represented as low electronegativity), which promotes the chemical reactivity of metals:

As electronegativity decreases, the reactivity of metals will generally increase due to an increasing power to donate electrons to more electronegative atoms.

This concept links students back to Module 1, where they reviewed the periodic trends with electronegativity and how this impacts the nature of bonds formed between atoms. Students can draw upon this learning to relate the position of the metal on the periodic table to its observed trends in reactivity and to evaluate any contradictory evidence they may uncover during this process.

#### REDOX reactions

Although every chemical reaction involves the transfer of electrons between reactants to produce products, not all reactions can be considered REDOX or reduction-oxidation reactions. The characteristic difference between the two is that a minimum of two reactants undergo a change in oxidation state. One reactant donates electrons and increases its oxidation number (by the number of electrons donated); another reactant accepts these electrons and decreases its oxidation state (by the corresponding number). This is a paired process; oxidation cannot occur without reduction and vice versa.

Mnemonics are useful tools to develop student understanding of what can be quite complex and confusing terminology:

* OILRIG – separated into the two components, OIL and RIG.
* OIL – Oxidation is the loss of electrons.
* RIG – Reduction is the gain of electrons.

Some additional key terms that may appear in various resources can include:

* An **oxidant** or oxidising agent is the reactant that **causes** oxidation in another reactant by itself undergoing reduction (the matching paired process)
* A **reductant** or reducing agent is the reactant that **causes** a reduction in another reactant by itself undergoing oxidation (the matching paired process)

##### Half-equations

Spectator ions are the reactants and products which do not participate in the REDOX reaction. They do not donate or accept electrons; their oxidation number remains constant.

Half equations are chemical reaction equations that show separately the oxidation and reduction equations to aid the investigation of the processes. The half equation shows each reactant and its corresponding product, including oxidation numbers, to demonstrate the change in the oxidation state. These equations uniquely include electrons as a reactant (for reduction half equations) or as a product (for oxidation half equations) and exclude spectator ions. For example, the reaction of magnesium with hydrochloric acid would normally be described as:

This is a REDOX reaction where magnesium increases its oxidation number (from zero to +2), and hydrogen decreases its oxidation number (from +1 to zero). Chlorine does not change its oxidation state (-1 to -1) and is, therefore, a spectator ion and excluded from the half and net ionic equations.

The oxidation of magnesium metal to form magnesium ions:

The reduction of hydrogen ions to form hydrogen gas:

Each half equation can then be combined to create the net ionic equation. Reaction stoichiometry is determined by combining the oxidation and reduction half equations to balance (and cancel out from the equation) the number of electrons donated and accepted:

Half equations can be derived from the [datasheet [PDF 880 KB]](https://educationstandards.nsw.edu.au/wps/wcm/connect/98664936-221f-4c49-88e1-d002ec69285c/chemistry-formulae-sheet-data-sheet-periodic-table-hsc-exams-2019.pdf?MOD=AJPERES&CVID=). Standard potential reactions are written as reduction half equations. Reversing the reduction half equation will yield the corresponding oxidation half equation.

#### Galvanic cells

A galvanic cell is produced where the reduction and oxidation reactions are physically separated into half cells to redirect the electron transfer through a connecting wire. The net flow of electrons (electric current) through the external circuit enables the chemical energy from the half-cell to be transformed into other forms in the circuit components.

The respective half equation represents the chemistry of each half-cell.

For a galvanic cell to function, each half-cell must contain an:

* Electrode – made of a metal that can oxidise, or an inert metal that provides a surface for electron transfer.
* The **anode** is the electrode in the oxidation half-cell.
* The **cathode** is the electrode in the reduction half-cell.
* Electrolyte – containing cations of a metal matching the electrode or which can reduce on an inert electrode.
* Anolyte is the electrolyte in the oxidation half-cell.
* Catholyte is the electrolyte in the reduction half-cell.

The electrode must contact the electrolyte for the galvanic cell to function.

Mnemonics are useful tools to develop student understanding of what can be quite complex and confusing terminology. Combined with the previous OILRIG mnemonic, students can use this to describe the structure and function of the galvanic cell:

* ANOXREDCAT – separated into the two components, ANOX and REDCAT.
* ANOX – Anode is where oxidation occurs.
* REDCAT – Reduction occurs at the cathode.

The two half-cells must then be connected to allow the exchange of electrons and ions:

* Electrons are transferred via a metallic conducting wire.
* Cations and anions are transferred via a salt bridge.

As the REDOX reaction proceeds, electrons are liberated from the anode (oxidation of the anode metal to metal ions which enter the anolyte), travel via the wire and into the cathode (reduction of the cations from the catholyte which accumulate on the surface of the cathode). This process causes a positive charge to develop in the anode half-cell (where electrons are lost and cation concentration increases due to oxidation of anode metal to anolyte metal ions) and a negative charge to develop in the cathode half-cell (where electrons are gained, and anion concentration decreases due to reduction of catholyte metal ions to cathode metal) which quickly retards the flow of electrons in the wire and the electrode reactions cease.

A salt bridge contains highly mobile cations and anions that flow in opposite directions to correct the charge imbalance produced by the net flow of electrons in the wire. The salt bridge must contact the electrolytes in both half-cells to function.

The salt bridge provides anions to the anolyte (drawn from the catholyte) to cancel the positive charge generated in the anode half-cell and cations to the catholyte (drawn from the anolyte) to cancel the negative charge generated in the cathode half-cell. This process completes the circuit of flowing charges, and the galvanic cell can continue to function.

For example, a galvanic cell consisting of a zinc and copper half-cell:

Figure 2 – a typical galvanic cell made from Zinc and Copper half-cells.

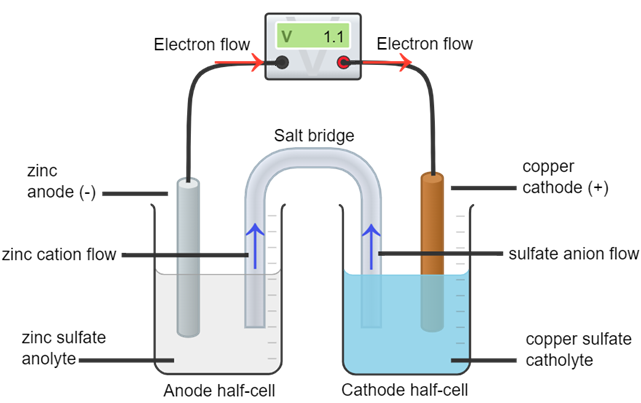


Image created using [Chemix](https://chemix.org/).

In Figure 2, labels indicate the direction of electron flow from the Zinc half-cell to the Copper half-cell and ion flow in the salt bridge.

These diagrams are typically represented with the electron flow running left to right. Therefore, this makes the left half-cell the anode and the right half-call the cathode.

Providing a blank template of these diagrams allows students to easily illustrate the chemistry and action of each half-cell and the function of the overall galvanic cell produced. Two versions of this blank template are provided in Appendix 1 for differentiation of this activity to develop student proficiency in describing galvanic cells.

Flow charts can assist students in tracing the chemical processes observed in the practical apparatus and the terminology and mnemonics used to describe their structure and function:

Figure 3 – flow chart of structure and function within a typical galvanic cell and the mnemonics to describe its function

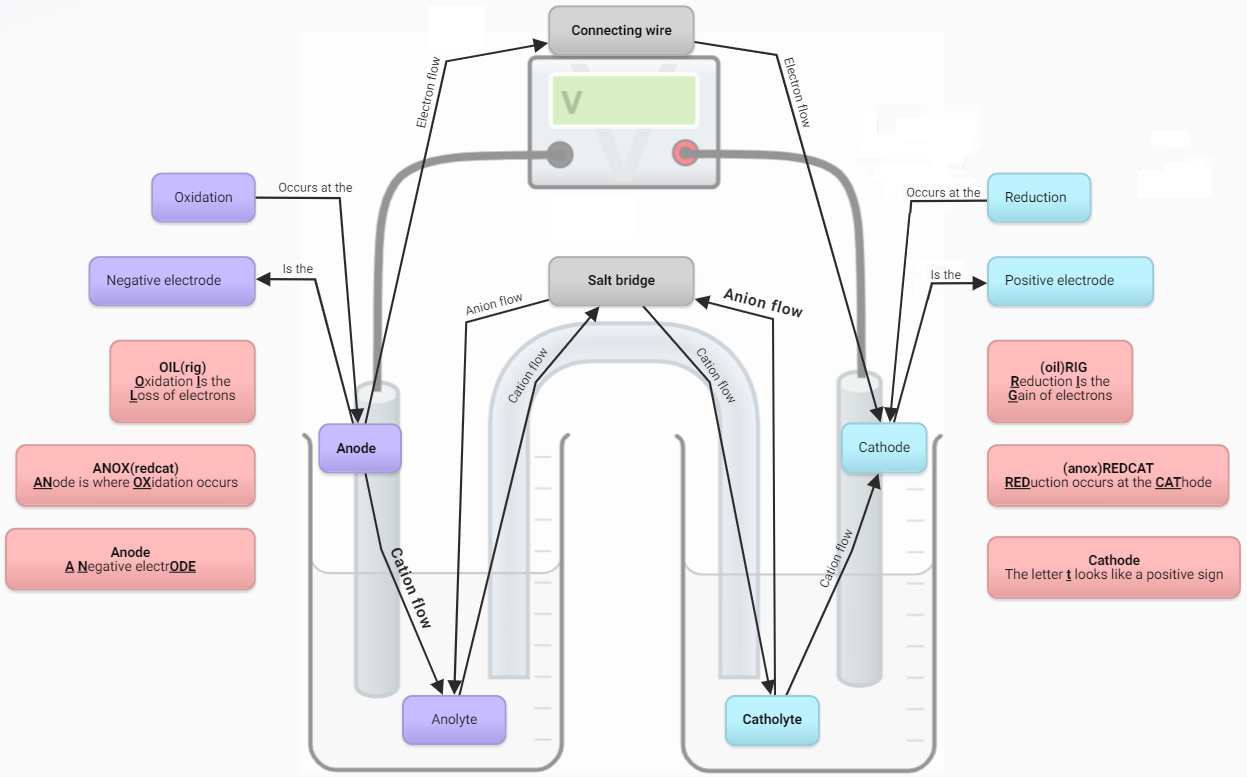


Image created using [Bubbl.us](https://bubbl.us/) and [Chemix](https://chemix.org/).

The flowchart in Figure 3 is superimposed over the diagram of a typical galvanic cell.

The galvanic cell will continue to operate while the following conditions (shown in bold in Figure 3) are met:

* The standard potential difference of the two half-cells is positive.
* The cell must have the moveable charge carriers:
* In the **connecting wire**, a path allows the flow of electrons from the anode to the cathode.
* In the **salt bridge**, a path for cations and anions allows cations to flow into the catholyte and anions into the anolyte.
* The electrolytes can provide ions to the salt bridge:
* In the catholyte, a source for anions, allowing the flow of anions into the salt bridge and onto the anolyte.
* the anode, a source for cations (in addition to the anolyte), allows cation flow into the salt bridge and onto the catholyte.

##### Cell notation

Shorthand notation of galvanic cells is commonly used to communicate the structure of the anode and cathode half-cells quickly and effectively. [Cell notation](https://chem.libretexts.org/Bookshelves/General_Chemistry/Book%3A_ChemPRIME_(Moore_et_al.)/17%3A_Electrochemical_Cells/17.09%3A_Cell_Notation_and_Conventions) can clearly show galvanic cells in a single line with the same structure as galvanic cell diagrams (anode half-cell on the left, cathode half-cell on the right):

Where the:

* single vertical lines represent the interface between the electrode and electrolyte in each half-cell.
* double vertical lines represent the salt bridge interface between the half-cells.

For example, the zinc and copper galvanic cell represented in Figure 2:

##### Standard potentials

Voltage is also known as the potential difference.

On the [datasheet [PDF 880 KB]](https://educationstandards.nsw.edu.au/wps/wcm/connect/98664936-221f-4c49-88e1-d002ec69285c/chemistry-formulae-sheet-data-sheet-periodic-table-hsc-exams-2019.pdf?MOD=AJPERES&CVID=), several standard potentials describe the reduction reactions of many metals, non-metals and some polyatomic ions. These potentials are the voltage measurements of each type of half-cell when connected to the [standard hydrogen electrode](https://chem.libretexts.org/Bookshelves/Introductory_Chemistry/Book%3A_Introductory_Chemistry_(CK-12)/23%3A_Electrochemistry/23.05%3A_Standard_Hydrogen_Electrode) (SHE) measured under [standard conditions](https://chem.libretexts.org/Bookshelves/General_Chemistry/Map%3A_Chemistry_-_The_Central_Science_(Brown_et_al.)/20%3A_Electrochemistry/20.4%3A_Cell_Potential_Under_Standard_Conditions). The reduction of hydrogen ions to hydrogen gas is set as zero volts in this table as the reference point for all other measurements. Each standard potential is shown as the **potential difference** to the SHE. The order of this list represents the relative capacity to undergo the reduction reaction as shown or the oxidation reaction by reversing the reaction shown:

* Negative reduction potentials correspond to half-cells, which are more likely to undergo oxidation – this species ‘feels negative’ about reduction.
* Positive reduction potentials correspond to half-cells, which are more likely to undergo reduction – this species ‘feels positive’ about reduction.

When connected in a galvanic cell, the half-cell with the more negative standard potential (higher on the standard potentials list) has a higher power to oxidise and will become the anode half-cell. This forces the other half-cell with the more positive (or even slightly less negative) standard potential (lower on the standard potentials list) to undergo reduction and become the cathode half-cell.

##### Cell potential

The cell potential (or Ecell) is the overall voltage generated between the anode and cathode half-cells connected in a galvanic cell. This voltage represents the maximum electromotive force the galvanic cell can provide based on the chemistry of each half-cell. The cell potential is calculated by:

This calculation takes the half-cell potentials for the galvanic cell and determines the **potential difference** between the two. The standard potentials taken from the data sheet are used without modification. The structure of the equation resolves any combination of negative or positive anode and cathode half-cell potentials.

Giving students a range of scenarios to practice these calculations is important to complete the discussion of galvanic cell structure and function:

|  |  |  |  |
| --- | --- | --- | --- |
| Cell notation | Anode half-cell | Cathode half-cell | Ecell calculation |
| Zn|Zn2+||Cu2+|Cu | Zn|Zn2+  -0.76V | Cu2+|Cu  0.34V |  |
| Mg|Mg2+||Zn2+|Zn | Mg|Mg2+  -2.36V | Zn|Zn2+  -0.76V |  |
| Cu|Cu2+||Ag+|Ag | Cu|Cu2+  0.34V | Ag+|Ag  0.80V |  |

The higher cell potential allows more electrical energy to be harnessed in the external circuit connected between the anode and cathode half-cells. This allows more work to be done by the electrons flowing through the circuit.

If students use a digital voltmeter to measure the cell voltage, the positive lead is connected to the cathode half-cell and the negative lead is connected to the anode half-cell. If the voltage reading obtained is negative, the negative and positive probes are connected in reverse to what is expected.

Additional questions can be explored to deepen student engagement with galvanic cells and relate the chemistry to their everyday use as batteries:

* What is the chemistry of a ‘flat’ battery? A battery would be considered ‘flat’ when it can no longer generate the voltage required for the continued function of the external circuit – but what is happening chemically?
* How and why does the structure of a typical battery differ from our galvanic cells?
* How do these galvanic cells differ from rechargeable batteries in modern devices?
* Why do we continue to use lead-acid cells in cars when less hazardous alternatives exist?

Presenting students with specific scenarios to challenge their understanding can also help to drive a deeper understanding of galvanic cells:

* Using [non-standard conditions](https://chem.libretexts.org/Bookshelves/General_Chemistry/Map%3A_Chemistry_-_The_Central_Science_(Brown_et_al.)/20%3A_Electrochemistry/20.6%3A_Cell_Potential_Under_Nonstandard_Conditions) – modification of temperature or concentrations of each half-cell and the impact this has on the cell potential. This can be related to the temperatures of the environments in which batteries are stored and used.
* [Concentration cells](https://chem.washington.edu/lecture-demos/concentration-cell) – both the anode and cathode half-cells have the same chemistry (a typical Ecell calculation would yield zero volts). The anolyte has a lower concentration than the catholyte. The cell will produce a voltage between these half-cells until the concentration of the electrolytes becomes equal.

### IQ3-3 Rates of reactions

#### Rates of reactions

The [rate of a chemical reaction](https://chem.libretexts.org/Bookshelves/Physical_and_Theoretical_Chemistry_Textbook_Maps/Supplemental_Modules_(Physical_and_Theoretical_Chemistry)/Kinetics/02%3A_Reaction_Rates/2.05%3A_Reaction_Rate) is defined as the rate of change between the reactants and products. When students consider chemical reactions, it is important to recognise the difference between those which occur virtually instantly (for example, combustion of methanol) and those which occur much slower (for example, rusting of iron).

Chemical reactions require collisions, but not every collision is successful. Successful collisions require reactants to collide with the correct orientation and sufficient energy. The rate of these successful collisions then determines the rate of the chemical reaction.

Students can complete several practical demonstrations and investigations for this inquiry question. Through observations of these chemical reactions, students can manipulate variables (where appropriate) to measure changes in reaction rate:

* Clock reactions, for example:
* [Blue bottle](https://edu.rsc.org/experiments/the-blue-bottle-experiment/729.article)
* [Traffic lights](https://edu.rsc.org/exhibition-chemistry/traffic-lights/2020080.article)
* [Iodine clock](https://edu.rsc.org/experiments/iodine-clock-reaction-demonstration-method/744.article)
* Oscillating reactions, for example:
* [Modified Belousov–](https://edu.rsc.org/experiments/a-simple-oscillating-reaction/753.article)Zhabotinsky reaction
* [Briggs-Rauscher reaction](https://chemistrytalk.org/briggs-rauscher-reaction/)

Online simulations and videos are available to support students in investigating the influence of different factors on the rates of chemical reactions:

* [PhET: Reactions and Rates](https://phet.colorado.edu/en/simulation/legacy/reactions-and-rates)
* [ChemCollective: Kinetics](http://chemcollective.org/kinetics)
* [How to speed up chemical reactions (and get a date) - Aaron Sams (4:55)](https://www.youtube.com/watch?v=OttRV5ykP7A)

It is important to emphasise to students that although several factors will increase or decrease reaction rate, there are limits to how fast or slow a reaction can proceed.

#### Temperature

Temperature is the measurement which represents the kinetic energy of molecules. The faster molecules move, the higher their measured temperature. Raising or lowering the temperature of a substance will increase or decrease the movement of the molecules within the substance. Increasing the temperature of a chemical reaction will typically increase the reaction rate by influencing the following factors:

* Activation energy – increasing the temperature will give reactants sufficient energy to satisfy the activation energy (Ea) ‘hump’ of the forward reaction.
* Collisions – increasing the temperature will give reactants additional opportunities to collide by the increased movement of molecules.
* Molecular orientation – increasing the temperature will allow reactants to collide with the correct alignment due to more frequent collisions.

Many possible investigations can be completed. Students can explore a range of scenarios from laboratory activities or real-world observations to develop an understanding of the relationship between temperature and reaction rate:

* [Glow sticks](https://chemdemos.uoregon.edu/demos/Lightstick-Reaction-Rates-versus-Temperature)
* [Sodium thiosulfate and hydrochloric acid](https://edu.rsc.org/resources/the-effect-of-temperature-on-reaction-rate/448.article)
* [Potassium iodate and sodium metabisulfite (starch indicator)](https://edu.rsc.org/resources/the-effect-of-concentration-and-temperature-on-reaction-rate/413.article)

#### Surface area of reactant(s)

A solid reactant's surface area will determine the space for potential collisions with other reactants. Grinding down a solid reactant into a fine powder increases the surface area and provides additional locations for successful collisions. Increasing the surface area of reactants of a chemical reaction will typically increase the reaction rate by influencing the following factors:

* Activation energy – the surface area does not change the activation energy of the forward reaction.
* Collisions – increasing the surface area will give reactants additional physical space for collisions to occur.
* Molecular orientation – increasing the surface area will give reactants additional opportunities to collide with the correct alignment due to more possible arrangements of the reactant molecules.

Many possible investigations can be completed. Students can explore a range of scenarios from laboratory activities or real-world observations to develop an understanding of the relationship between surface area and reaction rate:

* [Effervescent tablets](https://www.sciencebuddies.org/teacher-resources/lesson-plans/surface_area_reaction_rates)
* [Marble (calcium carbonate) chips and hydrochloric acid](https://www.chemguide.uk/14to16/rates/surface.html)
* [Rhubarb (oxalic acid) and potassium permanganate](https://edu.rsc.org/resources/rates-and-rhubarb/745.article)
* [Magnesium ribbon and hydrochloric acid](https://edu.rsc.org/experiments/the-rate-of-reaction-of-magnesium-with-hydrochloric-acid/1916.article)
* [Dust explosions modelling using milk powder](https://edu.rsc.org/resources/burning-milk-powder/830.article)

#### Concentration of reactant(s)

Concentration is the measure of the amount of a solute within a given volume of a solution.

Each molecule in a solution is capable of collisions. With increasing reactant concentrations (compared to solvent molecules), there will be an increasing frequency of collisions between reactant molecules. Increasing the concentration of reactants in a chemical reaction will typically increase the reaction rate by influencing the following factors:

* Activation energy – the concentration does not change the activation energy of the forward reaction.
* Collisions – increasing the concentration will give reactants additional opportunities for collisions to occur.
* Molecular orientation – increasing the concentration will give reactants additional opportunities to collide with the correct alignment due to more possible arrangements of the reactant molecules.

Many possible investigations can be completed. Students can explore a range of scenarios from laboratory activities or real-world observations to develop an understanding of the relationship between concentration and reaction rate:

* [Sodium thiosulfate and hydrochloric acid](https://edu.rsc.org/resources/the-effect-of-concentration-on-reaction-rate/743.article): this practical is also available from [Flinn Scientific](https://www.flinnsci.com/rate-of-reaction-of-sodium-thiosulfate-and-hydrochloric-acid/dc91860/) as a [video demonstration (7:01)](https://www.flinnsci.com/rate-of-reaction-of-sodium-thiosulfate-and-hydrochloric-acid2/vel1860/).
* [Marble (calcium carbonate) chips and hydrochloric acid](https://www.chemguide.uk/14to16/rates/concentration.html#top)
* [Potassium iodate and sodium metabisulfite (starch indicator)](https://edu.rsc.org/resources/the-effect-of-concentration-and-temperature-on-reaction-rate/413.article)
* [Rhubarb (oxalic acid) and potassium permanganate](https://edu.rsc.org/resources/rates-and-rhubarb/745.article)

#### Catalysts

Catalysts are further explored in Module 4.

Catalysts are substances that reduce the activation energy (Ea) ‘hump’ by providing a lower energy pathway for the forward reaction. Catalysts are not consumed in a chemical reaction. The forward reaction can occur faster under the same conditions by providing a lower energy pathway.

Providing a catalyst will increase reaction rates by:

* Activation energy – catalysts reduce the activation energy required for the forward reaction.
* Collisions – catalysts can act as temporary holders for reactants, allowing reactants to collide.
* Molecular orientation – catalysts will hold reactants in an orientation suitable for the reaction to occur. This promotes the favourable positioning of reactants, increasing the reaction rate.

Many possible investigations can be completed. Students can explore a range of scenarios from laboratory activities or real-world observations to develop an understanding of the relationship between concentration and reaction rate:

* [Decomposition of hydrogen peroxide with different catalysts](https://edu.rsc.org/resources/hydrogen-peroxide-decomposition-using-different-catalysts/831.article)
* [Sodium thiosulfate and iron (III) nitrate](https://edu.rsc.org/experiments/catalysis-of-a-sodium-thiosulfate-and-ironiii-nitrate-reaction/442.article)
* [Persulfate ions and iodide ions](https://www.chemguide.uk/14to16/rates/catalysts.html#top)
* [Sodium thiosulfate and hydrogen peroxide](https://edu.rsc.org/experiments/catalysing-the-reaction-of-sodium-thiosulfate-and-hydrogen-peroxide/1712.article)

## Appendix 1

Figure 4 – a blank galvanic cell diagram with labels

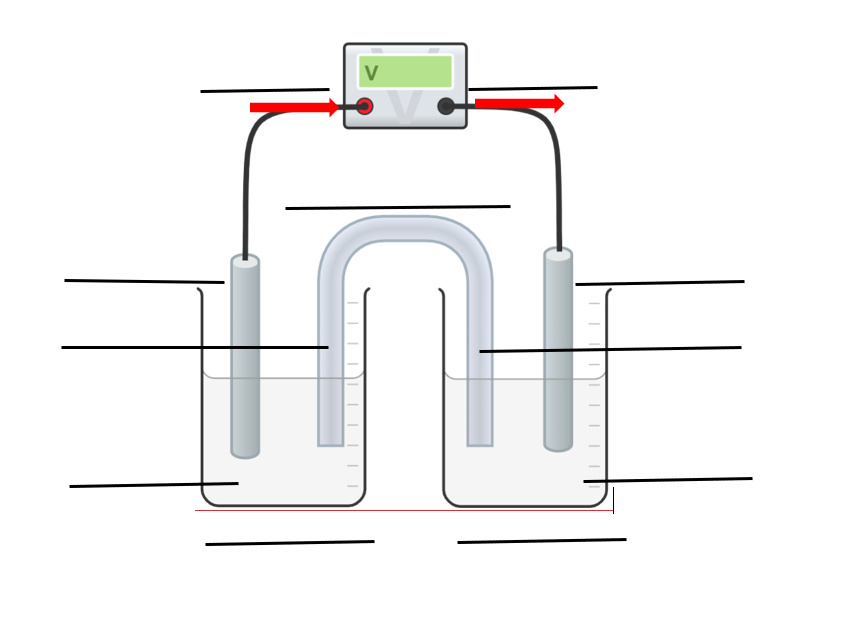


Image created using [Chemix](https://chemix.org/).

Figure 5 – a blank galvanic cell diagram without labels

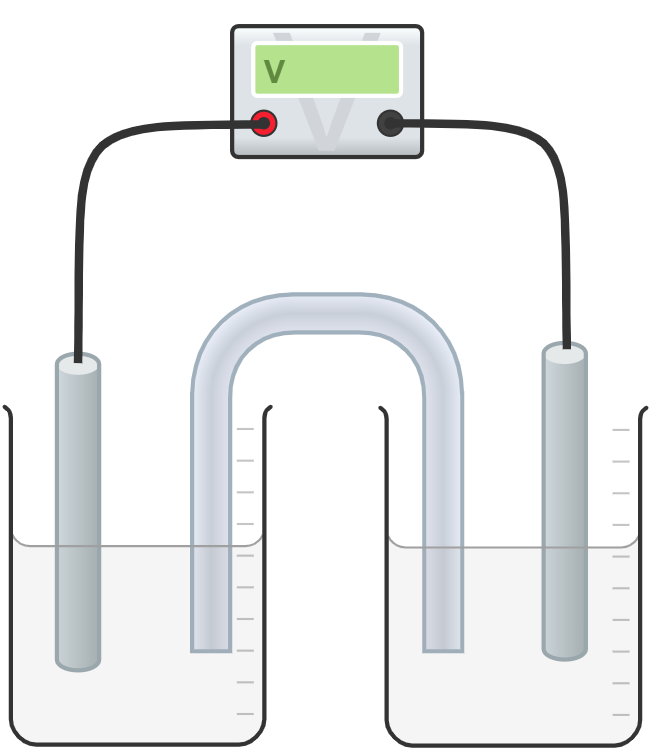


Image created using [Chemix](https://chemix.org/).

## Appendix 2

|  |  |
| --- | --- |
| Term | Definition |
| Chemical reaction | A reaction that can cause a change in the chemical bonding between substances. |
| Coefficient | It is the number before a chemical formula in a balanced chemical equation. |
| Detoxification | It is the process of removing toxic substances that can harm health or the environment. |
| Electrolytic cell | It is a cell that uses electrical energy to drive a chemical reaction. |
| Electrode | It is the site on the conducting surface (usually a metal) where electron transfer occurs. |
| Galvanic cell | It is a cell that makes a chemical reaction to produce electrical energy. |
| Gravimetric analysis | It is a quantitative method for accurately determining the amount of a substance in a mixture by selective separation. |
| Quantification | to express something as a number |
| Qualitative | Relating to, measuring, or measured by the quality of something. |
| Quantitative | Relating to, measuring, or measured by the quantity of something |
| Polyatomic ions. | These are cation or anion groups that comprise a combination of elements. |
| Redox | A reaction in which one substance is reduced and another is oxidised or deoxidised. |

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