



 **The Association
for Science Education**

Teaching Secondary Chemistry

Third Edition

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The principles behind secondary chemistry teaching

Vanessa Kind

Introduction

This book is written to help chemistry teachers enable their secondary school students, particularly those aged 11–16, to learn and understand the subject, connect ideas from disparate topics and enhance their interest in chemistry. The book contains eleven chapters including this one. Each of the remaining ten discusses how to teach a chemistry topic, such as particle theory, chemical change and acids and alkalis.

The book is one of a series of three Association for Science Education (ASE) handbooks; the others being parallel volumes in biology and physics. The first edition was published in 1999, followed by the second edition in 2012. This third edition retains the structure of previous editions but features new authors and substantially revised and updated chapters.

Throughout the book, the authors have kept in mind the needs of a secondary chemistry teacher confronted with the task of teaching a specific topic and the preparation required to deliver lessons that are satisfying educational experiences for teacher and students. Meeting the need for effective lessons that engage learners and enhance and sustain their curiosity lies at the heart of the book. Authors are aware that teachers approach lesson preparation with varied subject knowledge backgrounds: some have a deep understanding of chemistry, while others have little knowledge beyond their school experiences. Teachers with chemistry degrees may be specialists in sub-disciplines, with stronger subject knowledge of some topics than others. Our hope is that all teachers of secondary school chemistry, including those with extensive teaching experience, will find much of value in here.

This chapter examines the discipline of chemistry and discusses approaches to teaching that enable students to engage with the discipline, build their identity as chemists and learn conceptual ideas.

1.1 What is chemistry?

Chemistry is the scientific study of matter. Matter is anything that occupies space. Chemistry involves studying and understanding the behaviour of matter at micro- and nanoscale levels. This means understanding that all matter consists of a finite set of materials, the chemical elements, which themselves are constructed from tiny atoms and molecules. Many types of matter occur naturally. Others are developed in chemical reactions and processes devised through innovation, experiment and technology. Understanding and gaining control of factors such as temperature, pressure, rate of change and the amount of matter present is part of chemistry. Chemistry explains how matter is constructed, formed, changed and used; provides information about structures that matter adopts; informs measurement and control of the manufacture of new types of matter and discusses consequences arising from chemical processes. Part of the fascination of the subject is realising that chemistry is involved in almost every aspect of life on Earth.

A history of chemistry

Origins of the chemical elements

Chemistry begins with stars. The chemical elements, from which all matter is made, originate in stars. The simplest element, hydrogen, has been present in the Universe from the Big Bang onwards. Intense heat and pressure within a star cause atoms of hydrogen to collide with enormous force and fuse. Fusion creates atoms of many elements, including oxygen, nitrogen, carbon and iron. The heaviest atoms, including uranium, gold and platinum, form under extreme conditions in supernovae, which appear as bright, short-lived stars visible to the naked eye. In a supernova, the outer layers of an ageing star collapse inwards, creating a huge 'star quake' which expels matter into space at high velocities. In these conditions, heavy atoms collide and fuse, creating atoms that are heavier still.

The Earth formed about 4.5 billion years ago, most likely from a supernova, as a molten planet. On cooling, the Earth's surface solidified to a hard crust, retaining liquid surface water, a thin gaseous atmosphere and a molten core. Life began to evolve about 4 billion years ago and complex life about 500 million years ago. About 1.5 million years ago, *Homo erectus* was the first human ancestor to control fire, a man-made chemical combustion reaction. The ability to control fire led to chemical processes to make tools, weapons, plates, coins and jewellery from metals including copper,

gold, iron and bronze. By the fifteenth century CE, very high-quality steel called *tamahagane* was being made for Samurai swords in Japan; intricate figures and jewellery were being made from gold in South America; and in China, beautifully glazed ceramic jars and ornaments were manufactured. Wealth became associated with many of these items, particularly those made from precious elements such as gold and silver. The desire for wealth led to humans seeking methods for changing low-value base metals such as lead into rare 'noble' metals such as gold. Treating diseases that caused ill-health and death led humans to make medicines from plant and animal sources. An elixir was sought that would prolong life into eternity. These quests gave rise to alchemy, a proto-science practised from around 300 BCE onwards in Europe, Asia, the Middle East and on the Indian subcontinent. Alchemists devised techniques for measurement, heating and distilling. Their discoveries and practices laid the foundations for chemistry as a science based on observation and experiment.

Making discoveries in chemistry: products and techniques

In Europe, chemistry developed from this primitive state alongside huge changes wrought by the first industrial revolution that spanned the late-eighteenth to mid-nineteenth centuries. The invention and application of steam engines transformed small-scale local production of handmade items into large-scale mechanised processes, with factories manufacturing iron, fabrics, concrete, glass, paper and ceramics in huge quantities. These and other items were traded globally, contributing to enormous economic growth in many countries. The mid-nineteenth century brought a second industrial revolution in Europe, characterised by production of cheap steel of consistent quality made via the Bessemer process. The availability of steel transformed building and engineering, enabling construction of high-rise towers, longer and larger bridges and stable, long-distance railways. Discoveries of fossil fuels in multiple locations worldwide generated petroleum-based industries, leading to the development of the internal combustion engine and new materials based on crude oil. Huge expansion of chemical industries followed, as methods emerged for producing many new substances including fabric dyes, paints and pigments, fertilisers, explosives, soap, cosmetics, synthetic rubber and fabrics, pharmaceuticals, and foodstuffs including flavourings and colourants, factory-made sweets, chocolate and bread.

Throughout this era, chemists made fundamental discoveries that advanced understanding of the subject itself. Over half of the 92 naturally occurring chemical elements were isolated and identified in the nineteenth century as chemists mastered heating, reacting,

electrolysing, liquefying and measurement, and applied emerging analytical techniques such as spectroscopy, chromatography and X-ray diffraction. These prompted improvements in observations of phenomena and understanding of molecular structures. Theories were proposed and tested, and rigorous experimental practices that significantly improved understanding of chemical reactions were developed. Physical chemistry grew from work on the energetics involved in industrial processes. Inorganic chemistry arose from organisation and classification of the chemical elements via the periodic table and industrial processes developed for the production of acids, fertilisers, paints and other substances. Organic chemistry was founded on the petroleum industry and increasingly detailed knowledge gained from the study of living organisms. This led to chemists understanding the properties of carbon-based compounds, how they were formed, and how they could be adapted to make new consumer products.

Connecting to physics and mathematics

In the twentieth century, chemistry became increasingly sophisticated, incorporating discoveries and developments from physics and mathematics. Identification of sub-atomic particles and the application of quantum physics to electron arrangements led to new understandings of atomic and molecular structures. Chemists realised that molecules, atoms and ions combine via different types of chemical bond, so understood with greater precision how chemical and physical properties of substances are different and are related to the position of an element in the periodic table. Applications of mathematics enabled understanding of chemical equilibria, electrolysis, acid–base chemistry and molecular structures. Chemists applied newly developed techniques to chemical reactions in living organisms and biochemistry came to be recognised as the biological branch of chemistry.

New materials, new approaches

Within the last fifty years, chemists have applied understanding of the behaviour of matter to develop new substances including polymers, pharmaceuticals, fabrics, food additives and novel materials. Amongst the most significant developments are the computer chip, which uses the semiconductor properties of silicon, and revolutionised computing by enabling production of small-scale and hand-held machines; confirmation in the 1990s of the existence of a ball-shaped structure of 60 carbon atoms named buckminsterfullerene; and, in 2004, discovery of graphene, a one-atom thick sheet of carbon atoms arranged in an infinite network of hexagons. In pharmacology, new drugs – such as statins, anti-retroviral and cancer treatments – and

new vaccines have contributed to lengthening the lifespan of people in many countries. In chemical engineering, solar and photovoltaic cells, hydrogen fuel cells and battery-driven cars have been developed and are becoming commonplace as dependency on fossil fuels wanes in the light of evidence of climate change and as crude oil sources become depleted. Chemical processes have become greener, adapting to minimise waste, pollution and environmental impact. Analytical techniques and some reactions have been transformed by new technologies, including the scanning tunnelling microscope which allows chemists to 'see' atoms of chemical elements and perform reactions atom by atom.

Knowledge of chemistry is recognised as a foundation for understanding many subjects including medicine, materials science, geology, pharmacology, biochemistry and biology, forensics and chemical engineering. Chemical products contribute significantly to the economies of many countries: a recent report (Oxford Economics, 2019) suggests chemistry contributes around \$5.7 trillion dollars to global gross domestic products and supports 120 million jobs. Chemistry is a global enterprise. Studying and working in chemistry is productive and rewarding. Teaching chemistry means setting students on a path to understanding principles that underpin this industry and field of research, as well as how chemistry applies to their everyday lives.

1.2 Big ideas in chemistry

Two influential reports connect so-called 'big ideas' in science with the school science curriculum (Harlen *et al.*, 2010, 2015). Big ideas are aspects of science with one or more of these four qualities:

- Explanatory power; that is, understanding a big idea leads students to develop knowledge of a range of phenomena, objects and events.
- Provision of a basis for understanding societal issues.
- A sense of satisfaction, as big ideas prompt answers to questions people ask about the world.
- Cultural significance; that is, a big idea describes or reflects human achievement in science and how this impacts the environment.

In England, Northern Ireland and Wales, it is hoped that professional organisations for biology, chemistry and physics will work with curriculum developers to apply the reports in writing future versions of the science National Curriculum. For science teachers, these reports may facilitate departmental curriculum planning, ensuring coherence in students' experiences across age ranges, as the

example in Table 1.1 illustrates. The progression shown here extends to post-16 and indicates that effective and constructive 11–16 chemistry teaching builds students' understanding of concepts in a way which prepares them well for studying post-16 chemistry.

Often big ideas in science proposed by the 2010 report (Harlen *et al.*, 2010), two relate directly to chemistry:

- All matter in the Universe is made of very small particles.
- Chemical reactions involve rearrangements of the particles of reactants to form new substances.

Each is discussed briefly below. A third big idea, that the periodic table shows all known chemical elements in atomic-number order, is also introduced.

All matter in the Universe is made of very small particles

Understanding that all matter is made of discrete particles too small to be visible to the naked eye is the starting point for chemistry. This notion originated in ancient Greece when the philosopher Democritus (460–370 BCE) taught that matter was made of atoms ('atom' meaning 'indivisible') which were infinite in number, everlasting and uncreated. The Greeks thought matter could be divided down to atoms and no further. Democritus believed an object's properties derive from its atoms. Chemists now accept some of Democritus's principles as correct. Chemical elements sub-divide to atoms. This leads to a significant big idea in chemistry: that one atom of a chemical element has the chemical properties of that element. This means one atom of sodium has the chemical properties of sodium, one atom of platinum has the chemical properties of platinum, and so on for all chemical elements. This lays the foundation for understanding events occurring in chemical reactions.

Democritus's thinking remained current in the Middle East but was unfashionable in Europe until the seventeenth century, when scientists Isaac Newton and Robert Boyle stated their support for the concept. Formalisation of atomism in Western European science developed from 1808, when John Dalton (1766–1844) proposed his atomic theory. He emphasised that atoms could not be created or destroyed and that atoms combine to form chemical compounds with constant composition. Dalton reasoned that every water particle contains the same atoms combined in a fixed ratio. Dalton's initial principles were not entirely correct: at the end of the nineteenth century experiments

showed that atoms had sub-components, namely protons and neutrons in a nucleus surrounded by electrons. In the twentieth century, advances in understanding of the properties of electrons and mathematics led to the notion that electrons are arranged in 'shells' or 'orbitals'. Sub-atomic particles do not have the properties of individual chemical elements. Protons, electrons and neutrons are held in atoms by electrostatic and nuclear forces of attraction. Electrostatic attraction between particles of opposite charge is responsible for the chemical bonding that holds matter together and contributes to its physical and chemical properties.

Understanding that all matter, living and non-living, is formed from tiny discrete particles invisible to the human eye requires a mental leap from the observation of matter as continuous. Chemistry teachers must make atoms 'visible' to students. Chapter 3 discusses this in detail. Harlen *et al.*'s 2010 report illustrates how progression in students' understanding of the principle that matter is comprised of atoms can be developed (Table 1.1).

Table 1.1 Progression in understanding the big idea 'all matter in the Universe is made of very small particles'

5-7	All the 'stuff' encountered in everyday life, including air, water and different kinds of solid substances, is called matter because it has mass and therefore weight on Earth and takes up space. Different materials are recognisable by their properties, some of which are used to classify them as being in the solid, liquid or gas state.
7-11	When some substances are combined, they form a new substance (or substances) with properties that are different from the original ones. Other substances mix without changing permanently and may be separated again. At room temperature, some substances are in the solid state, some in the liquid state and some in the gas state. The state of many substances can be changed by heating or cooling them. The amount of matter does not change when a solid melts or a liquid evaporates.
11-14	If a substance could be divided into smaller and smaller pieces it would be found to be made of tiny particles. These particles are not in a substance; they are the substance. All the particles of a particular substance are the same and different from those of other substances. The particles are not static but move in random directions. The speed at which they move is the temperature of the material. Differences between substances in solid, liquid or gaseous states can be explained in terms of the speed and range of particle movement and the separation and strength of attraction between neighbouring particles. The stronger the bonds in between particles are, the more energy has to be transferred to the substance to separate them. Energy is transferred to a substance to become a liquid from a solid (melting) or to become a gas from a liquid. This is why substances have different melting and boiling points.
	All materials, anywhere in the Universe, living and non-living, are made of very large numbers of atoms. There are 118 different types of atoms known at the time of writing. Substances made of only one type of atom are called elements. Atoms of different elements combine together to form compounds. A chemical reaction involves rearrangement of atoms in the reacting substances to form new substances. The total amount of matter present in the reacting substances and the new substances remains the same. The properties of materials can be explained in terms of behaviour of the atoms and groups of atoms (molecules) of which they are made.
	Atoms have an internal structure comprising a heavy nucleus, made of protons and neutrons, surrounded by electrons arranged in orbitals or shells. Electrons and protons are respectively negatively and positively charged. Atoms are electrically neutral, as the numbers of electrons and protons present in an atom are the same, so their charges balance exactly. When one or more electrons are removed or added to an atom, an imbalance of electrical charge occurs and the charged particle is called an ion. Loss of one or more electrons creates a positively charged ion, because there are more protons than electrons present. Gain of one or more electrons creates a negatively charged ion because there are more electrons than protons present.
14-17	Electrons move rapidly in matter, causing electric current and magnetic force. The availability of electrons to move freely within a structure causes chemical elements to conduct electricity. A small number of chemical elements can be magnetised. Some atoms have nuclei which are unstable, due to an imbalance in the numbers of protons and neutrons present. These atoms spontaneously emit small particles in a process called radioactivity. This process involves the release of large amounts energy in the form of radiation. The behaviour of matter on the scale of nuclei, atoms and molecules differs from that observed on the scale that human senses can normally experience.

Source: Harlen *et al.*, 2010

Chemical reactions involve rearrangements of reactant particles to form new substances

Atoms, molecules and ions can combine to make new types of matter with properties that differ from those of the original reagents, whether these are chemical elements or compounds. An event in which a new type of matter forms is a *chemical reaction*. In any chemical reaction, the amount or mass of matter present does not alter regardless of the physical state of the reagents and products. Chemists cannot change the amount or mass of matter present, only the type of matter. This principle is the *law of conservation of mass*, which states that matter (or mass) cannot be created or destroyed in a chemical reaction.

Students may ask why chemical reactions occur spontaneously, that is, without the addition of any additional energy in the form of heat or a change in pressure. The answer lies in the overall energetics of the change from reactants to products. A reaction will occur spontaneously if the overall change of energy is positive. However, a positive energy change does not necessarily mean an increase in temperature, which is easy to identify. A positive energy change can mean that a gas is produced from two liquids reacting together, or by a liquid being formed when two solids react together. The positive change in these cases means that the products are more disordered than the reactants. The energy change may not be immediately measurable by a thermometer, but each individual particle of gaseous or liquid product has slightly more energy than those of the reagents. The total effect is an energy increase. Chemists use the term *entropy* for disorder: entropy always increases when a chemical reaction occurs.

Chemists have devised many methods for undertaking chemical reactions. To decide if a chemical reaction has occurred or is occurring, chemists make observations. Qualitatively, these include physical state, smell, overall appearance and colour. Quantitatively, chemists measure, for example, temperature, gas pressure, mass, concentration, colour changes and pH (acidity and alkalinity). Chemists try to understand why chemical reactions do or do not occur and how to make product(s) they want efficiently and sustainably. The ability to control chemical reactions is essential.

The periodic table shows all known chemical elements in atomic number order

The periodic table lists all known chemical elements, arranging these in an orderly way. The location of an element in the periodic table indicates its chemical and physical properties (see Chapter 4). Familiarity with the periodic table is therefore a valuable starting point for understanding chemistry. The periodic table originated from chemists' observations of patterns and trends in the physical and chemical properties of elements. Chemists considered how elements could be arranged logically. Early attempts included 'triads' of three elements and 'octaves' of eight. By the mid-nineteenth century, scientists found that many newly discovered elements would not fit into these simple arrangements. A comprehensive system that covered all the chemical elements was sought. In 1859, Dimitri Mendeleev, a Russian chemist, proposed the periodic table. He organised elements by atomic weight, stacking those with similar chemical properties above one another in columns called *groups*. Mendeleev aligned the elements into rows called *periods*. Enough was known about chemical elements to reveal that properties changed gradually, leading to the notion of periodic trends. Crucially, Mendeleev left gaps for chemical elements that were yet to be discovered. As these were discovered, Mendeleev's predictions were found to be correct.

1.3 Doing chemistry

Chemistry is a practical science. As a science, chemistry follows scientific principles. These include:

- For every effect there is one or more cause.
- Scientific explanations, theories and models are those that best fit the facts known at any one time.
- Knowledge scientists produce is used in some technologies to create products that meet human needs.
- Applications of science often have ethical, social, economic and political consequences.

In gaining understanding of chemical reactions and processes, chemists establish cause and effect. What chemists believe these causes and effects are has changed as scientific explanations based on techniques for measurement, analysis and observation have developed. In chemistry, application of these principles led

to paradigm shifts in chemists' thinking and understanding of the behaviour of matter. This deepened and refined knowledge about substances and how to apply chemistry to improve the quality of human life.

A paradigm shift: overthrow of the phlogiston theory

Until the late eighteenth century, many scientists believed that a substance named phlogiston was emitted from substances on burning. At the time, chemists lacked understanding of differences between gases; since the gaseous atmosphere that surrounded them was known as air, all gases were called 'airs' and named according to whether or not they supported combustion and other crudely observable properties such as solubility in water. 'Dephlogisticated air' supported combustion, while 'phlogisticated air' did not: a substance burning in dephlogisticated air gave off phlogiston, changing the air to phlogisticated air. The theory gained wide support and the supposed existence of phlogiston dominated chemical reasoning and analysis. This was despite some experimental findings that the behaviour of a number of different airs suggested otherwise. In the 1780s, experiments conducted by the French chemists Antoine Lavoisier and his wife Marie-Anne Paulze Lavoisier used very careful modern (for the time) measurement techniques and analysis to show that dephlogisticated air was a pure gas they called *oxygène* (oxygen). Phlogisticated air became carbon dioxide. The Lavoisiers effectively debunked the phlogiston theory, causing a paradigm shift that allowed chemistry to replace primitive reasoning with precise understanding. They developed systematic measurement techniques and chemical language based on elements present in substances. Their work laid the foundations of modern chemistry.

Chemistry that makes life better: a brief history of aspirin

Chemistry involves application of chemical reactions and processes to achieve human ends. As noted earlier, its origins lie in alchemy, in which people sought methods for manipulating natural substances to achieve personal benefits. Outstanding global progress in treating major diseases via vaccination, drug development and analysis arises because scientists, including chemists, have devised and applied technologies to solve challenges threatening human life. An example is the development of the drug aspirin, which has powerful pain-relieving, temperature-reducing and anti-inflammatory properties.

The history of aspirin begins with the willow tree (various species of *Salix*), a recorded source of medicine since antiquity: the Assyrian, Sumerian and Egyptian civilisations (4000–1500 BCE) used willow extracts to relieve joint pains, fever and inflammation; from around 600–500 BCE, Chinese and Babylonian herbal practitioners used willow as a remedy for various illnesses; the Greek physician Hippocrates (400 BCE) administered willow leaf tea to women in childbirth, while Dioscorides (100 BCE) found willow effective for treating inflammation. As societies developed, accurate knowledge of the active agents in herbal medicines was sought, as dosages and efficacy were inconsistent. By the mid-eighteenth century, medical and scientific societies enabled systemic recording and publication of experimental methods and results. In 1763, the Royal Society published five years of data on administering willow-bark extract gathered by an English country clergyman named Edward Stone. His precise preparation instructions and evidence that over fifty people were cured of fevers and other ailments led to willow-bark extract becoming a main drug available to herbalists. For researchers, Stone's work provided reliable information that a chemical in willow bark acted as a potent drug. Advancements in analysis enabled German chemist Johann Buchner to extract bitter-tasting yellow crystals from willow in 1828. He named the substance salicin after the Latin name for willow, *Salix*. Roughly simultaneously, salicin was extracted from meadowsweet flowers by the Swiss pharmacist Johann Pagenstecher.

Next, chemists investigated the structure of salicin. This was published in 1853 by French chemist Charles Frédéric Gerhardt, who named the compound salicylic acid, and synthesised a variant, acetylsalicylic acid. In 1876, systematic trials of salicin showed it was effective in treating fever and joint inflammation in rheumatism. The German chemist Felix Hoffman, who worked for the Bayer pharmaceutical company, found a method of making the drug in a laboratory. He also showed that the irritant side effects of salicin were reduced when acetylsalicylic acid was used, and he synthesised this molecule too. In 1899, Bayer named acetylsalicylic acid 'aspirin', coining the word by combining 'a' (for acetyl) with 'spir' (from *Spirea ulmaria*, the Latin name for meadowsweet) and 'in' (a common suffix for drugs at that time).

Aspirin became the leading drug for treating many illnesses and was awarded the title of 'most frequently sold painkiller' by the Guinness Book of Records in 1950. Further studies led to an understanding of aspirin's mechanism of action within the body. Today, aspirin is recommended for treatment of and as a preventative medicine reducing incidence of strokes, heart attacks and various forms of cancer.

1.4 Practical chemistry

The book introduces the practice of scientific inquiry as undertaken by chemists. This includes:

- Making observations and inferences based on real-time events.
- Identifying patterns in data and exceptions to patterns.
- Techniques for understanding substances and structures (chemical analysis).
- Techniques for making new substances (chemical synthesis).
- Examples of chemistry at work: the chemical industry and real-life chemists.

The nature of chemistry as a discipline means it is a hands-on, practical subject. While there are theoretical chemists, most people trained in chemistry gain experience in laboratories, manipulating equipment and chemicals. Experiencing chemical phenomena in a live setting generates motivation for teaching and learning theoretical ideas. In the UK, examination specifications for 14–16 and 16–18-year-olds require students to complete practical work, so hands-on experiences are essential. Chemistry experiments serve multiple purposes: they lead to knowledge and skill development, and require application of chemical principles, so creating a basis for advanced knowledge. Visualising reactions from virtual laboratories is no substitute for doing an experiment on the bench in real time. Even so, research evidence suggests practical work does not in itself teach chemical concepts, as students may follow instructions without knowing why they are doing so or the meaning of what they observe. Teachers should maximise the benefit (and expense) of practical work by giving careful consideration to expected learning outcomes.

Most chapters in this book feature chemical experiments, either teacher demonstrations or student-led activities, based on one or more of the principles of scientific inquiry listed above. Some activities produce immediately identifiable products that can be seen, smelt or tested qualitatively. Analysis skills progress to training students to make precise measurements of chemicals in, for example, titration reactions (Chapter 7). Other experiments introduce patterns of chemical behaviour displayed by sequences of elements in the periodic table (Chapter 4) or organic compounds with similar formulae (Chapter 10). Gradually, anomalies and exceptions to these patterns may be introduced. Post-16 students use advanced equipment and make a wide range of measurements, tests and observations. This methodology underpins chemists' development of productive and responsible applications. Open-ended, inquiry-based experiments raise questions that students investigate in a non-fixed

way, devising their own method. These permit various outcomes from one starting point, so introducing the principles of research and encouraging creativity.

1.5 Ethics in chemistry

Thankfully for the progress of science, and for the health and wellbeing of scientists, scandals involving lapses of ethics are relatively rare. Nonetheless, all scientists must adhere to ethical standards. They may be tempted to falsify data, perhaps in anticipation of a significant finding, so they can publish before a rival, gaining kudos and credibility. Disregarding health and safety regulations by habit or choice creates conditions ripe for serious accidents. Funds intended for research could be diverted elsewhere. Plagiarism transfers credit for work without acknowledgement. Junior members of a research team may be subjected to harassment, leading to stress. Working outside strict regulatory frameworks generates legal vulnerability. While all these are possible scenarios, in practice, the vast majority of scientists work reliably and honestly. For chemists, unexpected consequences of research probably cause the most significant ethical challenges. Thousands of new molecules are made annually, each part of a process of generating novel substances that may contribute positively to human development. However, some of these molecules may unintentionally have a negative impact on society or be deliberately turned to destructive uses. An example is discussed next.

Unintentional impact of a well-intentioned discovery: chlorofluorocarbons

Chlorofluorocarbons, or CFCs, were made innocently in a well-intentioned attempt to solve the problem of refrigeration. By the 1920s, electrically powered refrigerators were becoming a popular way to preserve food and other essential materials. Refrigeration relies on a refrigerant, a substance that changes state readily from liquid to gas and vice versa. Liquid refrigerant is pumped through pipes at the back of the refrigerator, absorbing heat from within the fridge and becoming a gas. As the gas leaves the refrigerator, a fan cools the pipes and the refrigerant re-condenses to liquid, transferring thermal energy to the atmosphere through a grille attached to the tubes. Initially, refrigerants were toxic and/or flammable substances such as ammonia, methane, propane, and sulfur dioxide. In the USA,

Frigidaire, a leading refrigerator manufacturer, sought a non-toxic, non-flammable alternative. Thomas Midgley Jr (1889–1944) obliged, synthesising the first CFC, dichlorodifluoromethane (CCl_2F_2), which was named freon. CFCs proved to be highly successful compounds ideally suited to their intended function. Midgley was feted for his contribution, winning international medals. By the 1970s, around one million tonnes of CFCs were produced annually for use as aerosol propellants, blowing agents for foams and packing materials, and as refrigerants. Their non-toxic nature led to CFCs being released into the atmosphere after use.

Meanwhile, in the 1960s, British scientist James Lovelock (1919–) noted build-up of atmospheric haze in certain weather conditions near his home. Curiosity led him to investigate the haze using his own invention, the Electron Capture Detector (ECD), a device sensitive to tiny amounts of synthetic chemicals. Lovelock detected CFCs even in places such as the west coast of Ireland and (in 1971) Antarctica where he expected to find 'clean' air. CFCs had dispersed across the globe. Lovelock presented his findings at a conference in 1972. His talk inspired American chemist, F. Sherwood (Sherry) Rowland to investigate atmospheric CFCs. With Mario Molina, Rowland published a seminal paper in 1974 showing that although CFC molecules are inert in the lower atmosphere, the troposphere, in the stratosphere, about 20 km above the surface of the Earth, interaction with UV-rays causes them to break up. Rowland and Molina calculated that a single chlorine atom from one CFC molecule initiates a chain reaction that destroys thousands of molecules of ozone, O_3 , the vital substance that protects the Earth from excessive UV-radiation. They estimated that an immediate ban on CFC production would mean ozone depletion continuing for years, while maintaining production would exacerbate the loss. Rowland said this realisation sent 'a chill down his back'. He became increasingly troubled as industry experts disputed his claims, in part because results were difficult to replicate in tropospheric conditions.

Independent evidence supporting Rowland and Molina came from British scientists Joe Farman and colleagues working for the British Antarctic Survey (BAS). BAS data from 1957–58 onwards showed annual depletion of stratospheric ozone during the Antarctic summer when the Sun reappears. They suspected something was destroying ozone. In 1985 Farman and his team published a paper demonstrating that ozone over Antarctica was depleted by 40% each September (the end of the southern hemisphere summer) and showed significant overall decline since the 1960s. The Antarctic ozone hole was quickly recognised as an environmental disaster, prompting

immediate change in human behaviour. By 1987, 56 countries had signed the Montreal Protocol, agreeing to halve CFC production, paving the way for a complete ban by 1996. Scientists expect Antarctic ozone to recover to 1980 levels by 2070.

1.6 Teaching and learning chemistry

Achieving conceptual change

Inevitably, students come to chemistry with naive knowledge which is often incorrect compared to formal chemical knowledge. The book gives many well-documented examples of naive knowledge in the section of each chapter entitled 'Students' prior knowledge and misconceptions'. Naive knowledge often (but not always) impedes learning of formal knowledge. Some naively held ideas are well-formulated into mental models that offer plausible explanations for phenomena and strongly resist change. It is worth noting that experienced scientists are as vulnerable to this failing as students. The phlogiston theory described above offers a good example: Joseph Priestley, a chemist often credited with discovering the gas that became known as oxygen, called it dephlogisticated air, and continued to believe in phlogiston despite strong evidence to the contrary. Similarly, students may persist in thinking that gases have no mass, solids 'disappear' on dissolving, matter is continuous not granular, and the periodic table is just a picture on the chemistry laboratory wall. Any learning that takes place is rote, for the convenience of passing an examination, and/or pleasing the teacher.

A major challenge for chemistry teachers is how to avoid being satisfied with students rote-learning formal chemical knowledge, instead helping them achieve deep understanding of at least the three big ideas discussed above. 'Deep' in this context means a major shift from naive to formal knowledge, not, for example, understanding quantum mechanics, mathematical details of entropy, or the intricacies of making new superheavy elements. Chemistry teachers must force students to abandon naive views, adopt chemists' ways of 'seeing' the particles comprising substances, realise what is meant by a chemical reaction, and understand the organising principles behind the periodic table and its impact on our understanding of matter. This formal knowledge creates an intellectually satisfying foundation for further study. The big question is how to achieve this knowledge and understanding.

Of course, this question cannot be fully answered in one paragraph. And it is not easily done. But a teacher who is serious about achieving more than surface learning can adopt four principles for their lessons.

The first principle is *not to make assumptions about what students know or understand*. Knowing what students know already is a powerful tool to achieve deep learning: establish students' prior knowledge, perhaps via a brief 'bell task' or introductory exercise (see White and Gunstone, 1992). Initially, this may feel like opening Pandora's box ('a source of great and unexpected troubles'), as realisation dawns that students' naive knowledge is complex, plausible, systematic and resistant to change. But students enjoy being asked to say what they think, and knowing their naive knowledge makes reconstruction to formal knowledge a possibility.

Second, *be prepared to ask hard questions*. The skill of asking and answering questions that probe understanding is essential. To prompt deep learning, find the question(s) that force(s) students to reconsider their naive views. These do not involve lazy description or guessing what is in the teacher's head. The questions must challenge long-held naive mental schemata. Finding the answers requires careful thought and time.

The third principle is to *provide support*. This means a variety of things. It includes continued reinforcement of new formal knowledge, such as confirmatory examples and situations that require application of the novel (to the students) reasoning. Support also means providing motivation for change, as students need reasons that are socially, attitudinally and emotionally acceptable in order to make the cognitive leap of accepting formal knowledge.

Finally, *take time*. This is often in short supply in science lessons: there is pressure to teach to a tight schedule, especially when examinations are looming. Faced with naive knowledge even after teaching, the temptation is to cut corners to formal knowledge, saying, 'You're supposed to know this' or, 'Don't think about it, just learn it.' Such grumpiness is tempting, but ultimately damaging to the process of achieving naive-formal knowledge shifts. It may seem hard to believe, but students whose formal knowledge is deep, not rote, have better examination results. Be patient.

Transferring between macro-, sub-micro- and symbolic scales: Johnstone's triangle

Chemistry explains macro-scale observable events in terms of sub-microscopic molecules, atoms, ions and electrons, representing them in chemists' language of symbols, formulae and equations. Chemists shift between macro, sub-microscopic and symbolic scales without

reflection, often assuming others follow their reasoning automatically. Students focus on the macro-scale, the level of observable events, and cannot make shifts between levels rapidly, in part because their understanding of the sub-micro and symbolic levels is not intuitive. A challenge for teachers is how to support students first, in understanding the sub-micro- and symbolic levels, and, next in moving between three levels of observation without access to extremely expensive equipment that enables them to 'see' atoms for real. A tool proposed by Scottish chemistry educator Alex Johnstone (1991) is valuable. Johnstone illustrates the macro-, sub-micro and symbolic levels using the triangle shown in Figure 1.1.

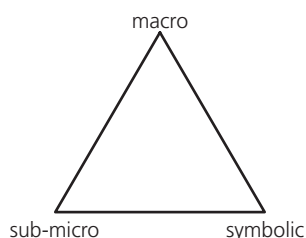


Figure 1.1 Johnstone's triangle

Source: Johnstone, 1991

The triangle scaffolds learning by prompting a need to be explicit when moving between macro-scale observations, sub-microscopic explanations and symbolic representations. The macro-scale is considered throughout the book when discussing observations and outcomes of experiments. Teaching students about particles, the sub-micro-scale, is considered in Chapter 3, and these ideas are applied throughout the book, beginning with modelling chemical reactions, discussed below. Symbolic representations of chemicals and reactions in formulae and equations is considered first in Chapter 2 and applied throughout the book.

Using models in chemistry teaching

Some molecular models have had enormous impact on scientific understanding. For example, Dorothy Hodgkin (1910–1994) deduced the structures of penicillin (in 1945) and vitamin B12 (in 1955), making models of both. In 1953, James Watson and Francis Crick elucidated the structure of DNA from photographs taken by Rosalind Franklin, building a three-dimensional model that changed understanding of genetics. In the 1980s, Harry Kroto made a paper model of C_{60} (carbon-60), based on an old football and futuristic buildings designed

by American architect Richard Buckminster Fuller, which changed our understanding of carbon allotropes.

Chemists have also developed theoretical models of chemical and physical phenomena. For example, attempts to understand acidity led to models of an acid, beginning with the seventeenth-century notion that acids were made of needle-shaped molecules, the sharpness of which caused their sharp taste. The twentieth-century Lewis model (the most recent) regards an acid as a lone-pair acceptor (see Chapter 7). Similarly, atoms were originally believed to be indivisible. In the late nineteenth century, J.J. Thomson proposed a 'plum-pudding' model in which negatively charged electrons (the plums) were embedded in a positively charged atom (the pudding); in the twentieth century this changed to a planetary motion model in which a positive nucleus is surrounded by electrons in specific orbits related to their energy; this was superseded by the quantum mechanical model in which orbits become regions surrounding a nucleus in which there is a probabilistic chance of finding an electron. Models representing phenomena are important because they illustrate the impact of empirical science on our understanding, show inaccuracies and prompt assimilation of new knowledge.

Students' understanding of the sub-micro scale can be aided by modelling chemical reactions using hand-held manipulatives, large-scale molecular models and animated visualisations. The purpose of modelling is to enable students to 'see' microscopic processes and form explanations for chemical events, rather than relying on macro-scale observations. These explanations consider, for example:

- the physical structures of reactants and products.
- the relative stability or level of energy that particles have before and after a reaction.

In chemistry at post-16 and beyond, models can represent reaction mechanisms, illustrating stepwise interactions between atoms, molecules and ions.

Initially, students benefit from seeing diagrams and images of molecules and manipulating molecules to track and trace a chemical reaction. These models support understanding of how products form from reactants, connecting the symbol equation for the reaction with observable macroscopic events. Placing images of molecular models alongside the equation and photographs of the reaction itself can help. For example, the reaction between methane and oxygen can be represented as shown in Figure 1.2.

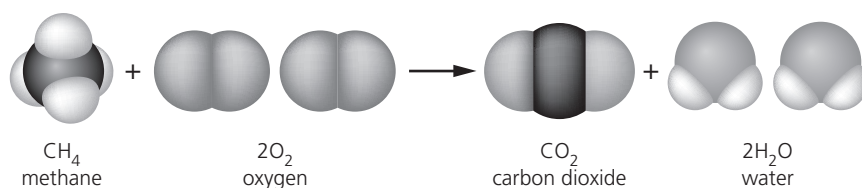


Figure 1.2 A pictorial representation to model a chemical reaction

In Figure 1.2, the image complements word and formula equations, showing the molecules involved. This model is space-filling, as no lines are drawn within molecules to represent chemical bonds. The shading shows that atomic cores are retained, as there are one carbon, four oxygen and four hydrogen atoms on both sides of the arrow. The model misleadingly gives the impression that atoms are genuinely unchanged: the carbon atom in methane 'relocates' and becomes the carbon atom in carbon dioxide. This is deceptive. Atoms are not full, intact atoms when bonded within molecules as electrons are shared across atomic boundaries. So it is important to make students aware this is a model, not reality.

Although models are extremely helpful in forming students' sub-micro reasoning related to chemical reactions, care is required, as students may develop misconceptions, some of which are listed below (misconceptions in *italics*).

- *Carbon atoms are black, hydrogen atoms are white, oxygen atoms are red, and so on.* In reality, atoms do not have these or any other colours. When observed in bulk, a substance may have a colour, but individual atoms are not coloured.
- *Atoms contain holes that need to be filled to make the atom 'happy'.* In practice, atoms do not have holes and are not solid. Nor do atoms have emotions so they cannot be happy or sad. Bonding occurs as a result of electrostatic attractions and energetics (see Chapter 6).
- *Bonds are physical objects, such as grey plastic sticks.* In reality, bonds are formed by electrostatic attractions between positive and negative particles (see Chapter 5).
- *Double bonds (bendy bonds) are made from two identical parts that are different from single bonds (straight bonds).* In practice, all bonds are formed by electrostatic attractions between particles.
- *Effort (energy) is required to make and break bonds.* Bond-breaking is endothermic, so requires energy, but bond-making is exothermic and emits energy to the environment. This can be shown when taking a model apart, as effort is needed. It is harder to use a physical model to show that energy is given out when bonds form.

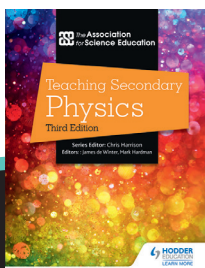
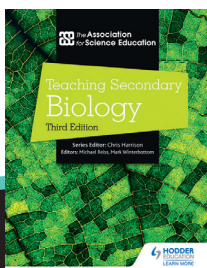
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