

Teaching Secondary Chemistry

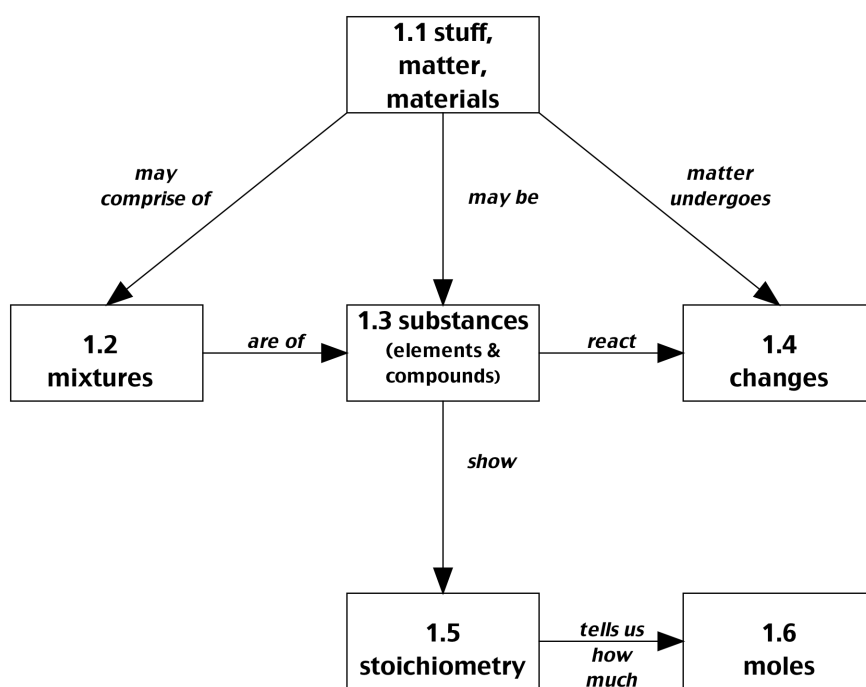
The version of record is:
Taber, K. S. (2012). Key concepts in chemistry. In K. S. Taber (Ed.), Teaching Secondary Chemistry (2nd ed., pp. 1-47). London: Hodder Education.

Chapter 1

Key concepts in chemistry

Keith S Taber

Topic Schematic



Choosing a route

This chapter will present key ideas that will be taught and developed through a spiral curriculum, with increasingly advanced treatments, throughout the secondary years – being revisited in different contexts. This is important, because many of the ideas met in chemistry are abstract, unfamiliar and even counter-intuitive. Learners therefore

need time to come to terms with these ideas – to explore them and become familiar with them. Few students will be able to master these ideas when first meeting them, so it is important that they are carefully introduced, and then later reviewed and reinforced in a variety of contexts. Luckily, chemistry as a subject supports this teaching approach, as many of the key ideas are directly relevant to teaching and learning across all topics. A balance will be needed between introducing new materials and revisiting previous teaching, as it is easy to overload students' working memories as they can only keep in mind a limited amount of new information at any time. However, effective and meaningful learning will require students to relate teaching to their developing understanding of the subject. The key is to recognise that whilst ideas are still novel (and often somewhat strange) they will place a demand on the learner, but if they are regularly reinforced in various contexts, then over time these increasingly familiar ideas will shift from being an additional load on memory, to acting as suitable support ('scaffolding') for new learning.

It is recommended, therefore, that after a major new idea is introduced (the distinction between single chemical substances and mixtures of substances, say), you should look for opportunities to review the idea as often as possible over the next few weeks and months. Initially treat the reviews as if dealing with new material (for some students they will be received that way), and over time shift to treating the ideas as taken-for-granted within the community of the chemistry class. Seeking regular formative feedback ('Jilly, can you remember what we called a substance with only one type of atom'; 'Vijay, could you remind the class what we mean by a chemical reaction?') will provide guidance on how quickly such shifts are possible with particular classes.

Similar advice is probably sensible in many subjects, but in chemistry we have to deal with two particular complications that do not always apply in other subjects. As some of our key ideas are abstract and not directly demonstrated, it is difficult to explain them clearly without reference to other equally abstract ideas. For example, consider the idea that a chemical change produces different chemical substances. To understand this statement, a student would already need to have a good grasp of the concept of chemical substances, so it would seem chemical substance needs to be introduced first. Yet understanding a chemical substance as something that retains its identity through phase changes (such as ice becoming water) to some extent requires

one to already have some notion that such changes are not considered as chemical changes. Of course, a decision has to be made about which ideas should be considered most suitable as a starting point for introducing students to these ideas, but students will not be in a position to fully appreciate these concepts when first introduced. So learning in chemistry is iterative, involving some ‘bootstrapping’ of partially understood concepts, one upon another.

Perhaps this circularity can be avoided to some extent by defining a pure substance somewhat differently – in terms of structure at the molecular level. Yet learning about ‘particle models’, as we will see in this and the subsequent chapters, is challenging for students, and we again run into the way ideas are intimately interlinked. Finding a simple ‘particle’ level definition of a pure substance that would apply unproblematically to all substances (e.g., neon, oxygen, water, common salt, sulfur, sugar and copper) might be a challenge for any teacher!

This chapter is then organised, as it needs to be, as a linear presentation of topics. In one sense this is a logical sequence to follow in teaching, as it does build up the complexity of the ideas. However, whilst I would advise teachers to try and follow something like this sequence, it is more important to realise that what order is chosen, the effective teaching of these ideas will not be achieved in a single pass through.

Previous knowledge and experience

Students will have had experience with materials as part of their primary education, as well as from their everyday experience of the world. Students are likely to be familiar in particular with the ideas of solids, liquids and (probably) gases, although their concepts here may be limited and imprecise (and talking about materials in this way, as being ‘solids’, ‘liquids’ and ‘gases’ may not be helpful, as will be explained in Chapter 2); and some may have met the particle model of matter, and may be familiar with simple representations of the states of matter at submicroscopic scale (though they are unlikely to have a good grasp of the actual scale at which these particles are considered to exist, nor a strong appreciation of the significance of the models).

1.1 Chemistry is about stuff

Definitions in science are notoriously unhelpful. They tend to either be very vague, too exclusive (i.e. seeming to omit things that should be included), or so technical that they are only useful to someone who already has a good understanding of what is being defined.

Chemistry is usually defined in terms of being about the nature, properties and structure of matter, or about the properties and interactions of different substances. Whilst not inaccurate, such definitions are of limited value to students until they have already started to see ‘matter’ in chemical terms, and to understand what chemists mean by ‘substances’.

Chemistry is about the stuff around us, and about thinking about this stuff in scientific terms. As a science, chemistry sets about analysing stuff in systematic ways, and this often means working with simplifications and generalisations – at least as starting points and ‘first approximations’. As a science, chemistry involves building up a body of theory: a collection of principles, laws, and models that can be used to make sense of, and so explain and predict, the properties of matter.

These are important points, as a student cannot be considered to understand chemistry in any depth unless she or he appreciates that as a science, its central ‘contents’ are not the phenomena in nature, but the theoretical constructs people have developed to explain those phenomena. Most chemists are very interested in the phenomena themselves – we tend to be fascinated by the colour changes, the ability to produce smells and bangs, and so forth. Students usually like this aspect of the subject: although for most the original fascination with smells and bangs is unlikely to last throughout secondary education if it is based purely on observing phenomena. What makes chemistry a science, and makes it a science that continues to fascinate many students, is the ability to organise and explain the phenomena in terms of models of great explanatory strength: models that with some modification can be applied across the wide range of substances and reactions met in school science (and of course beyond).

Substances

One of the major simplifications adopted in chemistry is to focus on substances. This is a simplification, because in our normal environment few of the materials we commonly come across are strictly substances in the chemical sense. Figure 1.1 sets out the relationship between some key ideas in chemistry. So where matter is a general term for stuff, we tend to use the term materials for well defined samples of stuff that we can work with – glass, wood, sodium carbonate (washing soda) poly(ethene), diamond, sea water, paint etc. From a technological perspective, these materials may have a similar status (different types of stuff that can be obtained, worked in various ways or used in different applications), but to a chemist they have rather different status. Materials may be pure substances like diamond or sodium carbonate, or mixtures such as air or paint.

This is a simplification as something like wood is more complicated, so that although it contains many substances, they are not simply mixed in a random way but built into a complex structure. Some manufactured materials are also composite, such as ‘fiberglass’, which contains fibres of glass embedded in a polymer (plastic).

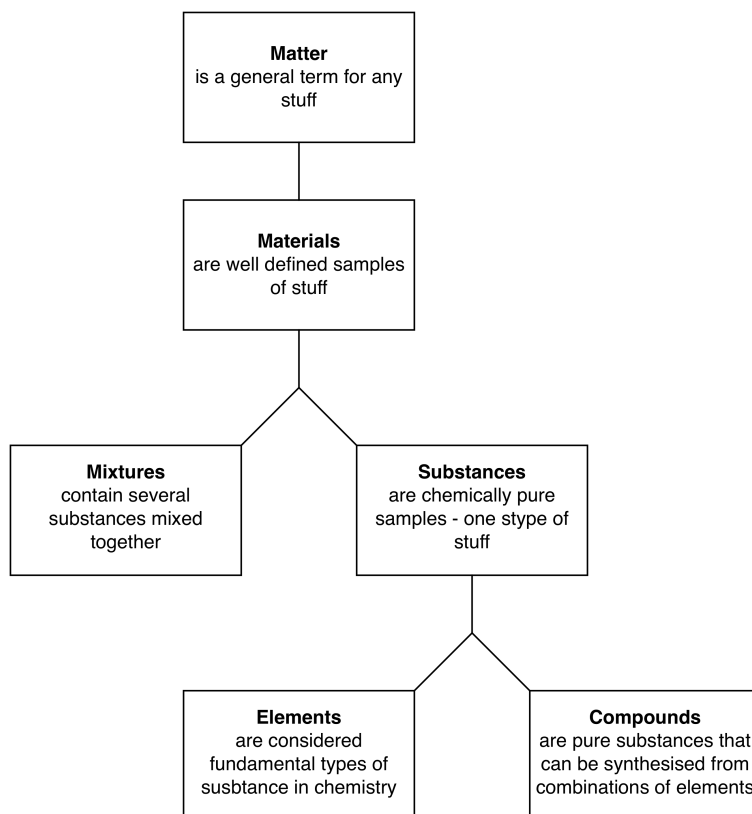


Figure 1.1: How some key terms are understood in chemistry

So most common materials in our environment such as air, sea water, earth, wood and even steel are not substances. This is a simple point but one which is not trivial for students. A key issue here in the minds of some students is the notion of ‘natural’ materials. For a chemist, natural products are those that derive from animal or vegetable sources, but are not considered to make up an intrinsically distinct type of stuff from other materials. (However, chemical terminology still retains vestiges of earlier thinking that living mater had some special vital essence, in our use of the terms ‘organic’ and ‘inorganic’.)

For many lay people ‘natural’ materials are considered to be intrinsically better (for example, safer) than ‘synthetic’ or man-made materials. The assumption seems to be that ‘nature’ knows best, and man less so. From a scientific perspective, man is part of nature, and any material that can be made by man is just as natural as anything secreted, excreted or extracted from a living organism. Indeed there are many berries,

fungi, insects, amphibians etc. that produce materials which are harmful or even lethal to people, whereas most synthetic products produced by chemists are subject to extensive safety testing before being allowed onto the market. Many natural products that were once difficult to obtain (e.g. from expensive processes to extract and purify tiny quantities of a substance present in living things) can now be synthesised much more effectively, and of course their chemical behaviour is unrelated to their origins.

As teachers, we need to be aware that many of our students may have absorbed at some level the notion of ‘natural-good, synthetic-bad’, and be prepared to challenge the supposition - without ignoring or underplaying how many synthetic materials can be used to do harm, in weapons for example, and may bring significant environmental costs in manufacture or disposal.

A closely related idea is that of purity. When buying orange juice to drink for example, we expect it to be ‘pure’ in the sense of just being material squeezed from oranges, and not including dead flies, sawdust, or the farmer’s finger-nail cuttings. To assure the potential buyer of this, the manufacturer may well claim to be selling “100% pure orange juice”, and in the context of selling and buying a drink this makes perfect sense.

However, students will need to be taught that no matter how pure our orange juice is in terms of only being juice from oranges, it is far from being a *pure substance* in chemistry! Orange juice is mostly water, but contains a wide range of other substances including fruit sugar, vitamin C, citric acid, various amino acids, and flavonoids that make orange taste different to lemon or grapefruit. Chemically, orange juice is a mixture of a lot of different substances, even though it is a natural product. A key distinction to be introduced and reiterated in teaching the subject, then, is that between *materials* which can be understood in everyday terms (orange juice is a different material to the glass, paper or ceramic cup we may drink it from), and the constituent *substances* that chemists analyse such materials into.

So the task of the chemistry teacher is to find a way to justify considering iron, but not steel; methane, but not petroleum; cellulose, but not wood; sucrose, but not honey; vitamin C, but not orange juice; and so on, as substances. As this is a difficult

distinction for those new to the subject, it is useful to have a wide range of examples that can be used when explaining the idea to the students.

However, the examples by themselves only seem persuasive to those of us who already appreciate the difference between materials in general, and those that are substances. Students will have to have good reasons to see this as a meaningful and important. There would seem to be two different approaches to thinking about what we mean by substances in chemistry, one of these is highly empirical, and the other more theoretical. Both approaches offer challenges for the teacher: but also a considerable opportunity to teach about the nature of science ('how science works').

An empirical view of substances

From a chemical perspective, materials are either pure samples of a single substance or consist of a mixture of substances. When a material is a mixture, it can in principle be separated into its components. There are a number of common separation techniques that can be used to separate different classes of mixtures, and a mixture that can be separated by one separation technique, will not necessarily be separated by another. For example, if sand is mixed with salt, it can be separated by dissolving (the salt) and filtering (see figure 1.2) - but that does not work if the sand is mixed with iron filings. Conversely, sand can be separated from iron filings by using a magnet: which does not have an effect on a sand/salt mixture.

[Figure 1.2. about here. Use figures 1.4 & 1.3 from first edition.]

Left-hand side: old figure 1.4

Add labels:-

To contents of beaker: Stirred mixture of sand, salt and water

To contents of basin: Salt solution

Add some material to inside of filter paper, label: Sand.

Right-hand side: old figure 1.3]

Figure 1.2 Separating sand and salt – salt is recovered by evaporation after dissolving and filtering

Laboratory exercises in this area can easily become somewhat artificial: for example giving students a deliberately prepared mixture of sand and salt for them to separate. That is somewhat different from being able to take an unknown material, and find out if it is a mixture, and – if so - separate it into its components. That was the kind of challenge faced by Marie and Pierre Curie when they carried out their work identifying new chemical substances (the elements radium and polonium). Marie Curie (see figure 1.3) has the distinction of being awarded Nobel prizes for her contributions to both Chemistry and Physics.



Marie Curie in her chemistry laboratory at the Radium Institute in France, April 1921.
Source: Nationaal Archief of the Netherlands
Photographer unknown
No known copyright restrictions

Figure 1.3 Image of Marie Curie in her laboratory

[This picture is taken from http://nobelprize.org/nobel_prizes/physics/laureates/1903/marie-curie-photo.html - note it says no copyright restrictions – this would need to be checked.]

It is often possible to recognise a mixture because mixtures usually do not have a distinct temperature at which they melt/freezing or boil/condense. Having identified a material as being a mixture, it is then possible to subject it to the battery of separation techniques available to the chemist. So for example, consider a liquid that was

considered to be a mixture (because a sample had been found to boil over a range of temperature). It may be that if the liquid is heated, one component of the mixture will evaporate, leaving a solid residue (for example, this would happen with salt solution). However, it is also possible that all of the mixture would boil-off. If a fractional distillation apparatus was set up (see figure 1.4) it may be that the different components have very different boiling temperatures, and that it is possible to separate them by carefully collecting condensate from vapour produced at different temperatures. However, it is also possible that the components could have similar boiling temperatures, making separation by this technique difficult as well. Perhaps another technique, such as a form of chromatography, might separate the components in such a case.

[Figure 1.4 to be based on figure 1.9, p.14 of first edition]

Figure 1.4 Fractional distillation to separate a mixture of liquids

A theoretical view of substances

From a theoretical perspective, a single substance is one that has a homogenous chemical composition: but as well as sounding like jargon, that seems a rather tautologous statement. The problem is that many mixtures, such as air, sea water, orange juice, bronze, etc, often *appear* uniform enough – we say they are homogeneous mixtures.

Once students have learnt about basic particle theory (see Chapter 2), and then progressed to learning about how many substances consist of molecules or ions (see Chapter 3), it becomes much easier to communicate the chemical notion of a pure substance. Distilled water, but not sea water, just consists of water molecules; sodium chloride consists of the same repeating pattern of sodium and chloride ions throughout the crystal; copper, consists of a repeating lattice of copper ions with associated electrons – whereas in brass, there are both copper and zinc ions present (but without a regular repeating pattern such as seen in the representation of NaCl structure in figure 1.5 below). It is clearly possible to show students diagrams of materials as chemists imagine them at the scale of molecules and ions to illustrate the distinction.

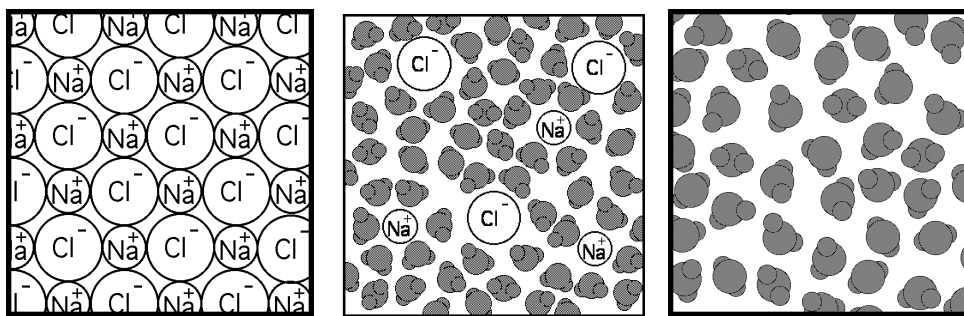


Figure 1.5: Two pure substances and a mixture

Figure 1.5 illustrates this point with the example of salt and water. To the teacher, the salient points to note about these images are likely to be clear – water is a pure substance because it only consists of one type of entity (water molecules, see the third image); sodium chloride is a pure substance as there is a constant composition (the alternation of Na⁺, Cl⁻ in the first image) throughout the material. The solution (the second image) shows the sodium and chloride ions mixed up with molecules of water. To the student, the distinction between pure substance and mixture may be less clear: the order in NaCl is not present in water (because it is in the liquid state, not the solid state), and - unlike water – sodium chloride does not comprise of one type of entity at this submicroscopic scale!

The images in Figure 1.5 are of course representations of models that we use in chemistry, and such representations offer challenge and opportunity. The challenge is in part that (as discussed below) molecules and ions are not quite like things we can draw as discretely bounded lumps; in part the difficulty of showing three dimensional, dynamic scenes in a flat image; and in part that in using introductory models we tend to need to ignore complications. So the third image in figure 1.5 shows that water only contains one type of particle, water molecules. Yet, actually, even pure water contains a very small proportion of hydrogen and hydroxyl ions, H⁺_(aq), OH⁻_(aq), which are important for some of its properties. Yet chemists do not consider that makes water a mixture rather than a pure substance, and it is not a complication we would wish to introduce when students first meet these ideas.

The opportunity here is to recognise that the limitations of our models actually reflect an important aspect of chemistry as a science. One of the key ways in which science works is through developing useful simplifications. It is not ‘wrong’ or ‘untrue’ to say

that water only contains water molecules: it is a useful ‘first-order’ simplification that is appropriate for many purposes. When students are encouraged to think about representations in this way, then the later introduction of increasingly sophisticated models more suitable for some purposes need not be seen as contradicting what students have previously learnt.

1.2 Mixtures

Previous knowledge and experience

Students will have much experience of mixtures of objects as well as examples of sets of the same type of thing in their everyday life, and so an approach building upon familiarity with such everyday examples is recommended.

A teaching approach

A good approach is to start with some models of mixtures at a level that all the students in the class can appreciate.

For example, some glass jam jars can be set up with various contents:

- A jar of steel ball bearings of the same size and appearance;
- A jar of marbles of the same size and appearance;
- A jar of marbles of different sizes and colours;
- A jar of marbles of the same size, but different colour inserts;
- A jar with a mixture of ball bearings, and marbles of a single size and appearance;
- A jar of ‘fruit and nut’ i.e. sultanas, currents, peanut, brazil nuts etc;
- A jar of mixed nuts;
- A jar of (unsalted and un-spiced) peanuts;

- A jar of salted peanuts;
- A jar of honey-coated banana chips.

Clearly, these are just suggestions, and the actual examples may be varied. You may wish to add additional examples, or in large classes have two or three jars of each example. The jars should be labelled with numbers or letters rather than descriptive names. Students should work in small groups (pairs or triads), and have to complete a table to show which jars they think contain mixtures. Importantly, students should be asked to agree in their groups, before recording an answer. Where the group does not initially agree, or is not sure, they should discuss the example, and try to come to an agreement. Warn students that they may be asked to give reasons for their decisions. The focus on dialogue is important both because research suggests such talk supports learning, and also because part of the role of the science teacher is to introduce learners to the nature of the ‘discourse’ of science: which involves examining evidence, presenting and discussing ideas, developing agreements, and seeking to persuade others of your ideas. Whilst the teacher and textbook are likely to be better informed about chemistry than the students, and part of teaching involves presenting and justifying currently accepted scientific models, too much authoritative teaching (the teacher telling, the students listening) gives a poor impression of what doing science is really like. Effective science teaching therefore involves a balance between developing students’ thinking and argumentation skills, and helping them learn accepted scientific ideas. This is especially important in chemistry where students will meet a progression of models during their time in school classes, as they need to appreciate that scientific models are human inventions that sometimes have limited ranges of application, and in some circumstances need to be replaced by more sophisticated thinking.

After the class has completed the exercise, hold a plenary session where you explore the decisions students made, and their reasoning (especially in any cases where groups do not agree). There is clearly room for some disagreement about what counts as being similar enough not to be considered a mixture. If students have not noticed that the peanuts in one jar are salted, point this out, and ask if it makes a difference?

In the case of the banana chips point out that the contents are not just banana, as these are honey-coated banana chips. Ask students whether that makes a difference to their decision? I suggest this latter example because of an episode when the author saw a label in a ‘health food’ shop for ‘a mixture of honey-coated banana chips’. Being a science teacher, I asked the shop assistant what the honey-coated banana chips were mixed with. The shop assistant did not understand the basis of the question. I concluded that the shop was selling honey-coated banana chips mixed with other honey-coated banana chips – not a good model of a mixture to a chemist. This point is worth emphasizing to the class – when they meet particle diagrams of compounds, it is quite common for students to think that if there is more than one element represented they are dealing with a mixture, even if there is only one type of unit (molecule) present (see figure 1.5).

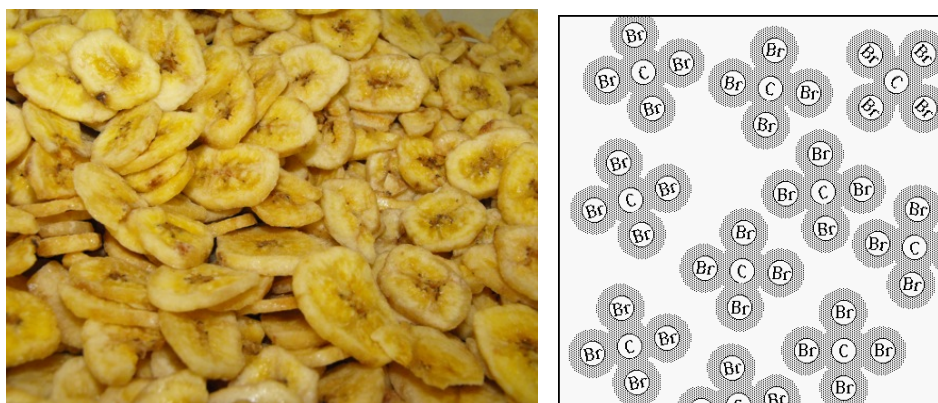


Figure 1.5: Mixtures? Both the honey-coated banana chips and the compound are best understood as having one basic type of unit.

[\[\(Note: photograph is taken from web – need replacement image with copyright permission\)\]](#)

For those jars that are considered to contain mixtures, ask students whether they could be changed into something that was not a mixture. This would mean using a separation technique. In these examples, this could be quite crude, as tweezers or even fingers could do the job. You might want to start demonstrating this process just to reinforce the idea. With some groups you might decide to select a couple of students who like to be doing things with their hands to separate the marbles and ball bearings

into two different jars. Asking them to do this by transferring between jars with tweezers might add a little challenge! With other groups, simply making sure they recognise that such a separation is possible will suffice. Modelling a process such as filtering would be possible for some mixtures, using a suitable gauge garden sieve.

The next activity moves from a model, to actually considering materials, so use boiling tubes sealed with bungs or corks to indicate that you are moving from a model to looking at the kind of examples of interest to chemists. Students are again asked to judge which samples are mixtures, but this time they are asked to consider examples such as:

- Sand mixed with salt;
- Sugar mixed with salt;
- Salt solution;
- Black ink;
- Air;
- Iron filings mixed with sulfur powder (flowers of sulphur);
- Iron filings mixed with copper turnings;
- Water;

There is the opportunity here for moving to a range of practical work that will allow students experience of using basic laboratory equipment.

For any practical activity (including teacher demonstrations) the teacher should:

- Check on the process for risk assessment that is required in the school, and either identify and follow a suitable existing protocol, or undertake a new risk assessment for any novel activity;
- Refer to the latest regulations regarding substance hazards, to ensure that students are not put at unnecessary risk. For most practicals there are suitable low hazard options.
- If in doubt refer to a more experienced colleague – if then still in doubt, the practical should not be undertaken. In the

UK, the organisation 'CLEAPSS' provides an excellent information service to schools and colleges, and most schools are subscribed (usually through the local authority) and have key publications available in the school preparation ('prep') rooms;

- Remember that assessing risk is not just about the activity, but also the people and the conditions. The same practical may be viable with some teaching groups and not others; and class size and the specific teaching room (e.g. arrangement of furniture, gas taps and sinks) should be considered.

As always with practical work in science, it is important to ensure students' minds are working as well as their hands – and so they are thinking about the scientific ideas behind the activities. This means differentiating activities for particular classes, or even for particular groups within classes. So ideas would be:

Dissolving and recrystallising

Showing that different salts have different solubilities in water is a useful exercise. Students can be asked to modify a simple 'add solid to water, stir, and observe' activity to increase the challenge:

- Can they provide a quantitative measure of solubility?
- Can they see if water temperature effects solubility?
- Can they see if the amount of material dissolved changes over time?

A given amount of solvent will dissolve a specific amount of a particular solute, at a certain temperature – often, though not always, solubility increases as temperature increases. However, it may sometimes take a noticeable time for a solution to become 'saturated'.

Recrystallisation is used as a purification technique. Here the challenge may be to make crystals from an initially fine grain (powder) sample, by using a dropping pipette to add samples of solutions to a watch glass or Petri dish, and waiting for

evaporation. Using hand-lenses or low-power microscopes can add to the excitement of making crystals.

Separating salt from sand

This activity (see figure 1.2, earlier in the Chapter) involves a sequence of steps: adding water to the mixture; stirring; filtering; evaporation of the solvent from the solution. When the author used this activity with first year secondary students, a DART (Directed Activity Related to Text) was employed to ensure students thought about the logic of the sequence of steps. The DART consisted of a set of labelled diagrams showing the steps and their end-states. Groups had to cut out the images, and sequence them correctly before starting the practical work. A similar process could be completed as a whole class activity on an interactive white board, with the final agreed sequence left on display while the class work.

Chromatography

The most advanced forms of chromatography are very 'high tech', with gas-liquid chromatograms costing tens of thousands of GB Pounds. However, separating the dyes in coloured inks using filter paper is a traditional school practical (see figure 1.7) that most students enjoy, and which shows that we cannot always tell what is a mixture simply by inspection.

Figure 1.7: Paper chromatography

[Use figures 1.11 and 1.13 from first edition

Figure 1.11 on left, add a second patch of substance for sample C, level with patch produced for sample A; figure 1.1.3 on right.]

Good separation of coloured inks is usually possible with alcohol (e.g. butan-1-ol) based solvents, but these are flammable and require good ventilation. Good results can *sometimes* be obtained using water as the solvent, which is more suitable for younger students. However, this does mean finding suitable pens to use in advance, as many permanent inks will not be appropriate.

Teachers can set up a ‘forensic’ context for practicals such as this: with a sample of ink from a crime scene, and samples from several suspects. Whilst this may seem artificial, students often enjoy identifying the offender, and it is a good example of how chemistry is used in forensic science.

1.3 Elements and compounds

The ancients referred to air, earth, water (and fire and the ether) as elements because they considered these substances basic. We still hear the air we breathe and the water we drink referred to as ‘elements’ (for example, in expensive advertising campaigns), as this seems to tap into something in the popular psyche. However, in chemistry, the term ‘element’ is reserved for a limited number of substances. Referring back to figure 1.1, there were two major distinctions shown there, between materials that were mixtures, and others that were substances; and then between substances that are elements and those that are compounds. This is potentially a very difficult idea for students to grasp – if substances are pure, single types of stuff, then how can some of them be compounded from others? Figure 1.8 revisits Figure 1.1, but acknowledges this issue. The distinction beneath the dashed line is likely to make more sense to students once they have learnt about atoms, ions and molecules.

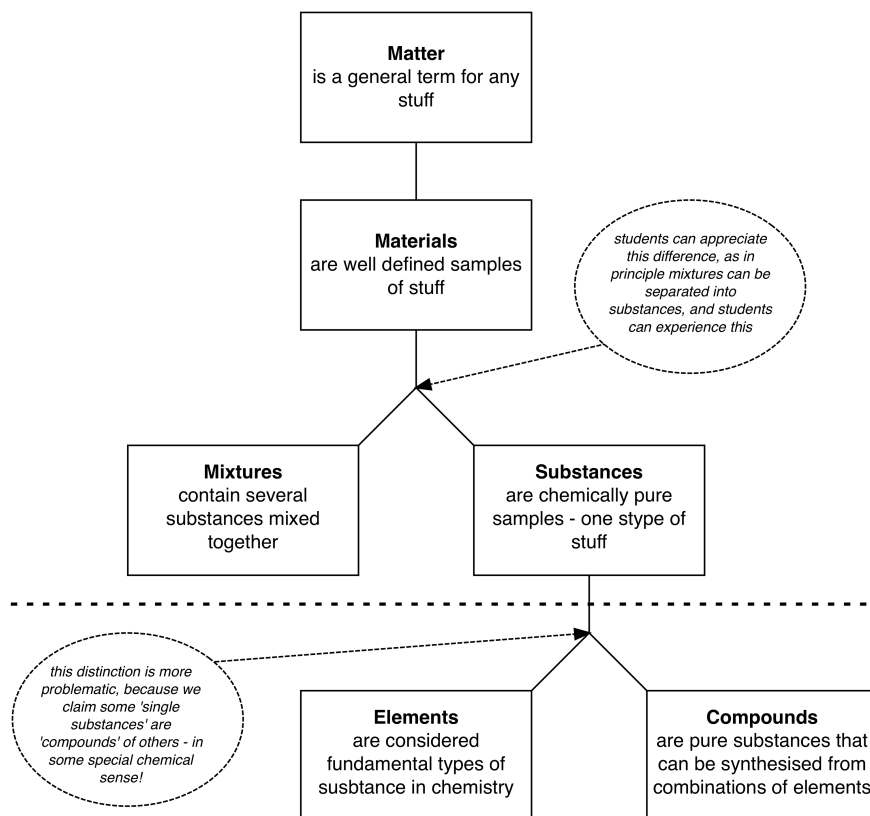


Figure 1.8: The distinction between elements and compounds is important, but may be challenging for students

Indeed the dashed line in figure 1.8 highlights one of the reasons that learners find the element concept a challenge. The modern scientific notion of element is a kind of hybrid concept, containing at least three somewhat distinct facets. This is shown in figure 1.9, represented as three overlapping meanings.

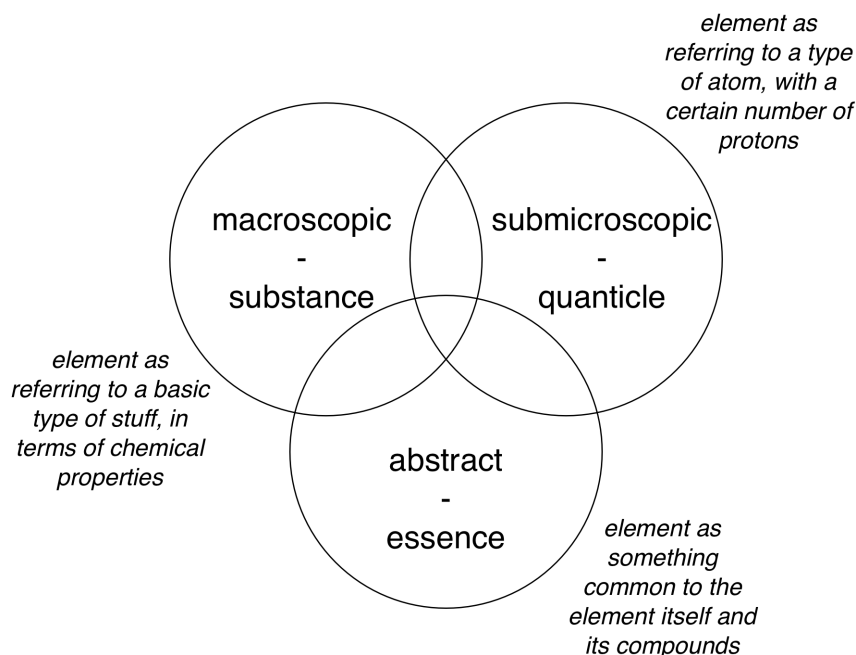


Figure 1.9: Three facets of the ‘element’ concept

So here we are talking about an element as a basic kind of substance. In terms of chemical properties, it is one that (unlike a compound) cannot be converted to any more basic substance by chemical means. However, the periodic table of the elements is often presented in terms of properties of atoms – and in particular the distinction in terms of atomic number, i.e. the number of protons in a nucleus defining the element (hydrogen: 1; helium: 2; lithium: 3; etc). We are working here with two complementary meanings for the idea of element, one at the (macroscopic) level of phenomena we can demonstrate to students (substances, and their reactions); the other deriving from a theoretical model in terms of conjectured submicroscopic entities (‘quanticles’, see below, §1.5).

However, there is also a sense in which an element is considered to be present, in a virtual or potential sense, within its compounds. This use is more common among French-speaking chemists, and in the English-speaking world we normally consider it quite inappropriate to suggest that sodium is somehow present in sodium chloride, or hydrogen in water. Yet, of course, chemical formulae (NaCl , H_2O , etc) tell us that the compounds somehow ‘contain’ the elements.

These distinct meanings of the same term are unfortunate, and can lead to confusion among students. For an element (as a substance) has a set of properties that would distinguish it from either a collection of its atoms, or from one of its compounds: so the substance sodium has quite distinct properties from a large number of discrete sodium atoms, as well as having different properties to the other substances that are its compounds: sodium oxide, sodium chloride, etc.

An additional complication, and one that arguably is best put aside when first teaching about element, is the existence of allotropes: where samples of the same element (in the term of the atoms present) can exist in different physical forms with different properties. So, for example, solid samples of carbon can exist as diamond or graphite (or indeed other forms) at the same temperature and pressure. They have very different physical properties (hardness, electrical conductivity). As another example, common oxygen (dioxygen, O₂) and ozone (O₃) are both gaseous forms of the element oxygen (both only contain atomic nuclei with 8 protons), yet they have different chemical properties.

Diamond and graphite are certainly different materials, although both forms of the same element, carbon, and so are *generally* considered to be the same substance. Yet if dioxygen and ozone have distinct chemical properties, it could be argued that are actually different substances, although both forms of element number 8, oxygen. This shows the limitations in the set of concepts and demarcation that chemists use to describe and classify the different substances found in nature. This complication is unhelpful when introducing the key ideas of substance and element, but when it is raised (by a particularly astute student) it offers a context for discussing an important aspect of the nature of science. Our chemical concepts – element, substance, acid, oxidizing agent, etc – are human constructions designed to help us make sense of the patterns we find in nature. Some of these human constructions fit less problematically upon the patterns we observe than others – but nature is often too subtle and nuanced for our simplest classifications and definitions to always work.

Teaching about elements and compounds

The distinction between elements and compounds is one of the most fundamental ones in chemistry, so it is important that students are able to appreciate it. However, there is no immediate way to distinguish elements and compounds, as both are single substances. Pure samples of either elements or compounds may be reactive or inert; and will give the sharp pattern of phase transitions expected of pure substances. That is to say, unlike a mixture, they will change state at a sharp melting/boiling temperature. So, unlike the distinction between single substances and mixtures, there is no simple way of demonstrating whether a substance is an element or a compound.

Having some sealed tubes with labelled samples of different elements and compounds can illustrate this well. Avoid materials such as alkali metals and phosphorous that must be kept in a protective medium, which could confuse younger students.

The common definition that an element cannot be changed into anything simpler by chemical means needs to be treated carefully. It is correct, as long as we consider an element to be ‘chemically simpler’ than a compound – but to a student that may seem a rather circular argument. The notion that there are some substances that are more basic, and which can combine to give all the other substances, is fine in principle. Yet the historical development of this area, shows just how much evidence and argument was needed to establish our modern understanding of the elements.

Most students will accept the basic principles here, and there may be a temptation to simply present the idea of chemical elements as if it is unproblematic. However, it is probably better that students appreciate something of just how much hard work was involved in establishing this basic idea: that the careful and difficult experiments and measurements of a good many scientists over a long period of time led slowly to our modern understanding. If time allows, it may be worth showing some pictures of early chemists at work in their (often make-shift) laboratories, to show that much that we take for granted in science today was once cutting-edge, and the basis of intense debate. This can give students a better feel for what working in science is like than simply presenting the outcomes of previous scientific work as in the sanitised form in which ideas are often reported in textbooks.

This is again an area that will benefit from an iterative treatment during the secondary years. Once students have progressed to learning about atoms (see Chapters 2 and 3), it becomes possible to shift to a definition of an element as a substance with only one kind of atomic core ('kernel') present. This ignores isotopes, of course, and so this is an area where it is useful if students have started to develop a feel for how we are using models in chemistry:

Way of thinking about elements	Comment
Most basic kinds of substance	Could only be useful if we have a criterion for what makes one substance more basic than another
A substance that cannot be broken down into anything simpler by chemical means	Again, this would only be useful if we need to already know what we mean by simpler
A substance that contains/is made up from only one type of atom	Only useful, once students know about atoms. Few elements (the noble gases) actually contain discrete atoms in their structure. Does not acknowledge isotopes.
A substance that contains only one type of atomic core	Acknowledges that most elements exist with (covalent, metallic) bonding that means they do not contain atoms as such. This is a more complex idea, though. Does not acknowledge isotopes.
A substance that contains only one type of atomic nucleus	Avoids the issues of whether there are atoms in most elements. However, students need to be aware that the nuclei are balanced by electrons in the structure. Does not acknowledge isotopes.
A substance that contains atoms/nuclei with the same number of protons	This definition allows for isotopes, but is more abstract than referring to 'one type of atom'.

Table 1.1: Complications of defining an element

Table 1.1 presents some of the advantages and disadvantages of different ways of defining the notion of an element. Clearly a parallel list could be drawn up for compounds. The most useful approach to take will clearly depend to some extent on the particular class, and in particular on *how much prior knowledge* they bring. Ideas

based on notions of atoms, or subatomic particles, can clearly only be introduced *after* students have learnt about atoms and atomic structure respectively.

The existence of isotopes, versions of an atom with the same number of protons but a different number of neutrons, adds a level of complication, in that different isotopes have atoms/nuclei that are in this sense different, but distinguished by a property which is not considered central to defining the element. Generally, samples of an element will contain a mixture of different isotopes, and so the relative atomic mass (the mass of an atom of the element relative to a standard, usually an atom of carbon-12, i.e. with 12 nucleons) quoted for element usually reflects this. For example, the relative molecular mass of chlorine is about 35.5 reflecting a roughly 3:1 mixture of chlorine-35 (17 protons, 18 neutrons) and chlorine-37 (17 protons, 20 neutrons). Isotopes (like allotropes, discussed above) represent complications that teachers need to decide *when* to introduce, and this may mean initially leaving them aside until students are confident with the more basic ideas.

However, it should be recognised that even the explanations which do not draw upon atomic ideas are abstract. So to consider a compound as a *single* substance which can be broken down into more basic substances is going to be challenging for many learners. A possible teaching model here might be a jar of peanuts still in their shells (i.e. groundnuts). This can represent a pure substance as the jar only contains one kind of object. However, it is possible to process the nuts to break up these objects and separate them into kernel and shells. However, care is needed in using such an analogy to ensure that students appreciate the kernels and (now broken) shells *represent* elements that were joined into a compound. Both the kernel and shell are actually made of complex materials and students need to appreciate they are being used as components of a model.

The chemist's elemental analyser

Of course chemists do have a device that can determine whether a substance is an element or a compound, and this is the mass spectrometer. The mass spectrometer will break up a molecule (whether of a single substance, or from a mixture) to produce discrete atomic ions, which can be detected by their different masses (or

technically, the mass to charge ratio, but most ions produced have unit charge). In many applications of mass spectroscopy, molecules in a sample are broken into various fragments that collectively give clues to the overall structure. However, if the sample is treated so that it is fully atomised, then the spectrum produced shows the range and relative numbers of different atoms present. So atomic mass spectroscopy acts as an elemental analyser!

The theory of the technique is beyond most secondary age students and is a 'sixth-form' topic, but the existence of a machine that can decompose a compound and show that at some level it contains more than one kind of component, could be useful when first introducing the idea of elements and compounds. Indeed, just as with a technique like chromatography (see above), it can be used to illustrate the possibility of separation long before students are ready to appreciate *how* it works (see Figure 1.10).

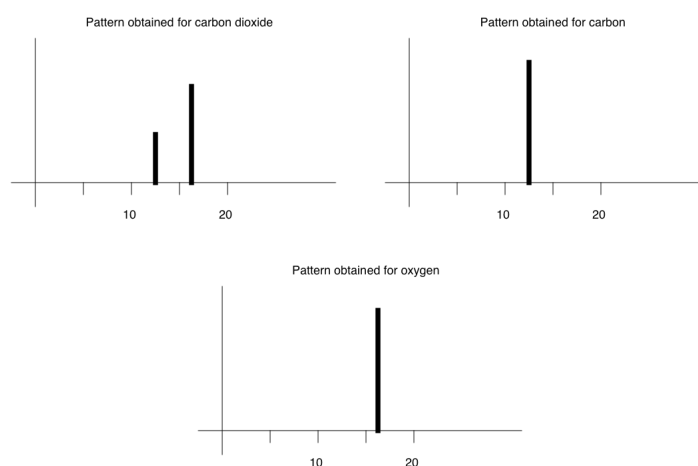


Figure 1.10: Chemists have a technique which helps identify elements and compounds – carbon dioxide is a single substance, but can be shown to ‘contain’ the elements carbon and oxygen

I would suggest that although the black-box approach to using mass spectroscopy from early in secondary school need not be problematic in itself, it is important for students to realise we are talking about something here which is not just another technique that separates the components of mixtures. Rather it should be seen as *chemically* changing a single substance that is a compound into its elements. Mass spectroscopy relies on physical separation techniques using electric and magnetic fields, but before this can happen a sample has to be *chemically* decomposed. So mass

spectroscopy is a technique that decomposes compounds (a chemical step) so that they can then be physically separated.

Of course, figure 1.10 again ignores the issue of isotopes. There is nothing wrong in using this kind of simplification, but I would advise telling students that there are some complications which are being ignored, because they only become important when chemists need to look at things in more detail. It is useful to seek to make a teaching point by being explicit whenever you use a simplification or model, both because it avoids students feeling they have been misled later, and because it is important for students to realise that science proceeds through the development and testing of various models, representations and theories. This is 'how science works'.

This point can be revisited later in the school, once atomic structure and isotopes are studied, when more complex diagrams showing isotopic composition can be considered. Even at that point, there is no need for students to know the details of how the technique works, although the basic physics required for a qualitative understanding is usually taught at upper secondary level.

The great advantage of talking to students about mass spectrometry is that it provides a basis for accepting the idea of elements as something special, and for identifying when we are dealing with them. This can allow the teacher to set students simple practical work that can be then discussed in terms of elements and compounds without the circularity noted above.

For example, the electrolysis of water, using the Hoffman voltameter apparatus, could be followed by presenting diagrams of what the mass spectra of the water and the gaseous products would look like – what chemists would find if they tested the three different substances (the water, and the gases collected at each electrode) in their 'elemental analyser' (see figure 1.11).

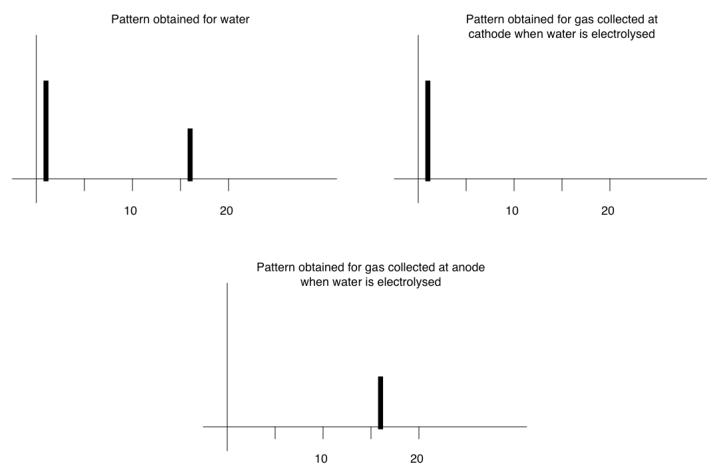


Figure 1.11: Electrolysis of water converts a compound into two elements

As another example, consider the common practical where weighed magnesium is heated in air (usually in a crucible to minimise the loss of the powdery product, and to avoid exposure to the bright visible and ultraviolet radiation emitted in the reaction) to demonstrate that the product has a greater mass than the reacting magnesium. The practical is often included in a topic on burning, sometimes as an illustration of counter-evidence to the phlogiston theory – the historical idea that burning is the release of something (called phlogiston) found in flammable materials, and responsible for that property. This was a focus of major debate in the development of modern chemical ideas, and is sometimes used as an example of how science proceeds (‘how science works’).

Combustion is a reaction with oxygen, and common examples of burning lead to an *apparent* loss of material (as the products are often formed as gases such as carbon dioxide), but magnesium forms a solid oxide. Students could carry out this practical work (you are advised to have some suitable sample results available, as inexperienced hands may well find the weighed product less massive than the magnesium they started with).

In a plenary session, students’ ideas on what is going on in the experiment, and what the measurements may be taken as evidence for, can be invited. Before closing down discussion, and explaining current scientific thinking (which should include making sure students realise that measured changes in mass are due to not weighing

everything present before and after a reaction), the class can be shown the results chemists get when they test magnesium, and the powder obtained by burning magnesium, in their elemental analyser (see figure 1.12).

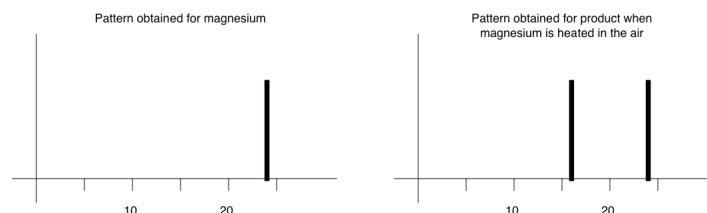


Figure 1.12: The product from burning magnesium includes a second element

If the idea of the elemental analyser is already familiar from earlier work, then students should be able to suggest from figure 1.12 that:

- magnesium is an element – one type of substance;
- the powder produced in the reaction is made up of two basic substances,
- one of which is the original magnesium

You can also tell students that the powder has been found to be a single substance, as it has a precise melting temperature, so as it ‘contains’ two elements it must be a compound.

The elemental analyser is a flexible idea that can be introduced in many contexts to reinforce the distinction between elements and compounds. It can help us with the rather abstract notion of how a compound can be said to contain elements. This can be a major problem for students. Sodium is a dangerous metal, and chlorine is a nasty gas that was used to kill people in war. If sodium chloride is a compound that contains sodium and chlorine, then surely – many students think - it should exhibit some of their horrible properties. The sodium in the compound should react with water, and the chlorine should attack the respiratory systems of those who come in contact with it.

This would be true if sodium chloride was *a mixture*, but a compound is a single substance that has its own properties quite unlike its ‘constituent’ elements. Sodium chloride does not contain *the substances* sodium and chlorine, and so they are not present to exert their properties. Only if sodium chloride is subjected to the extreme conditions of the elemental analyser can it be broken down to show that in a sense it is ‘made of’ the elements sodium and chlorine. Too much ‘salt’ is bad for our blood pressure, but we will not be badly burnt or choked by it, and a little is important in our diet.

The periodic table of the elements

This discussion of the sense in which a compound contains its elements reflects the ambiguity in the core chemical concept of the ‘element’ (see Figure 1.9). Chemists use the term ‘element’ to refer to both the basic substances themselves, and something more abstract that might be thought of as their ‘essence’, and which can be considered to be present in both samples of the elemental substance, and its compounds. What is actually – physically - considered to be present in both these contexts is nuclei with particular proton numbers – again something that can only be discussed with students after they have learnt about atomic structure.

When chemists were developing an understanding of the elements; and attempting to identify them, and distinguish them from compounds (from which they were sometimes very difficult to extract), they realised that one thing that seemed to distinguish elements from each other was how much of one element would react with another a certain amount of another element. By considering mass ratios involved in various reactions, chemists slowly started assigning masses to the different elements. It became clear that they varied widely, but that there were patterns in how much of one element would react with a certain amount of another (see ‘stoichiometry’ below), that were easier to understand once masses were assigned to the different atoms so that the ratio of atoms could be considered.

It was eventually recognised that the relative masses of the atoms was not the crucial feature, but something that varied along almost the same sequence as masses: what is now called the atomic number. Figure 1.13 shows a modern arrangement of elements

in the periodic table, showing both the atomic numbers, and the approximate relative masses of the elements.

Figure 1.13: A modern version of the periodic table of elements

[Use figure 7.5, p.225 from the first edition]

The periodic table is iconic, and there are many variations available. Indeed a quick web-search shows adoptions of the periodic table format to systematise desserts, rock bands and cartoon characters! More significantly it can be appreciated that there is no single correct form of the period table, but rather that it is a type of model used to organise chemical ideas and information – and so different versions are especially useful for different purposes. This again reflects the nature of chemistry as a science: scientists develop models to help them understand aspects of nature. Different versions of the periodic table can best reflect different aspects of the patterns among the elements that chemists have discovered in nature. For example, some versions of the periodic table are more aligned with the chemical properties of the elements, whereas others are organised according to the electronic structures of the atoms of the elements. That the same *basic* arrangement of periods and groups fits both of these very different considerations reflects how well atomic theory helps explain chemical processes observed in the laboratory.

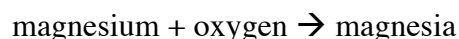
The background to the periodicity of the elements can be demonstrated by plotting charts of various properties against atomic number. This can be done by hand, but could also be a useful ICT-based activity, where data from a spreadsheet can be used to produce various types of chart. An important teaching point is the relative merits of using line-graphs compared with bar charts. As there is no meaning to interpolating between elements (there cannot be an element between element 15 and element 16 in the way there can be a time between 15s and 16s for example), line-graphs formats are useful for highlight *trends* – such as ionisation energy changes down a group – as long as it appreciated that atomic number is not a continuous variable.

This type of activity has traditionally been reserved for more advanced levels of study (where the patterns are considered in terms of underlying theories, for example to explain patterns in ionisation energies). However, such an activity could be useful as a

means of linking this topic with the nature of chemistry as a science (what is sometimes referred to as ‘how science works’). For example, a class of upper secondary students could be divided into a number of groups each of which is given the task of plotting a different property against atomic number, and considering whether there is any evidence for considering that property to repeat in a periodic pattern. Different groups could make brief presentations, before a class discussion to synthesise ideas. Such an activity could allow students to practice scientific argumentation, and focus on the relationship between ideas and evidence in science. Unlike many school activities, the evidence in this case is complex, and does not easily lead to a clear conclusion, showing what an intellectual achievement the original development of the periodic table was for chemists who did not yet know about atomic structure.

1.4 Change in chemistry – the concept of reactions

Chemical changes, also known as reactions, bring about a change of substance. Consider the example discussed earlier of magnesium burning in the air:



In this process we start off with magnesium. This is in chemical terms a single substance. It is a metal that has a shiny appearance when clean, conducts electricity, melts at 650°C, and burns with a bright white flame. Air is a mixture of gases, but the reactive component that is present in quite a high concentration is oxygen. Oxygen is a colourless, odourless, gas that can be liquefied (at normal pressures) below about -219°C (i.e., its boiling temperature is 54.36 K). After the magnesium has burned in air, there is no magnesium left. If we carried out the reaction in a sealed container, with just the right amount of oxygen present to react with the amount of magnesium used (see the next section on stoichiometry), then there would also be no oxygen left afterwards. These two substances have ‘disappeared’ in the reaction. However, a new substance has been produced. Magnesia (chemical name, magnesium oxide) is a white powder, quite different from either magnesium or oxygen. It is an insulator used as a refractory material (when high temperatures are needed), as its melting temperature is 2852 °C (3125 K), and is it hygroscopic – that is, it will absorb water from the

atmosphere. It is also used as an indigestion remedy. The new substance, the product, was not present before the reaction, and the reactant substances no longer exist. However, the total amount of material has not changed (3g of magnesium will react with 2g of oxygen to give 5g of magnesia). The elemental analyser (see above, figure 1.14) shows us that the original substances *are in a sense* present in the product.

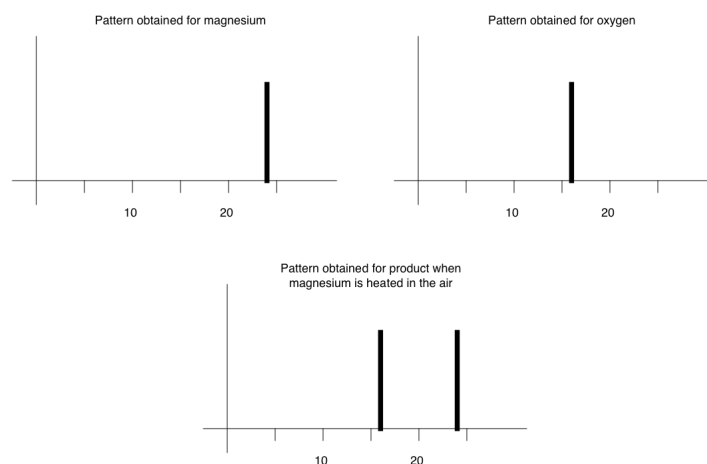
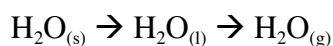


Figure 1.14: Evidence for a chemical change

Chemical and physical change – a useful rule of thumb?

A distinction that is often introduced in school chemistry is between physical and chemical changes. After a chemical change we have a different substance or substances than before. After a physical change we have the same substance in a different state or phase. So if ice is warmed it will melt, and if the water obtained is heated, it will boil to give steam:



Now ice, water and steam have some very different properties, and can be considered different materials. However, scientifically they are different states of the *same* chemical substance: hydrogen oxide (or, rather undemocratically, just ‘water’). These changes - ice melting, water boiling - are not chemical changes (see Chapter 2). Yet, to a novice student, such changes may seem just as dramatic as some chemical

reactions. So appearances do not always give us a clear indication of whether a change is chemical or not.

Traditionally a number of criteria have been used to contrast chemical and physical changes (see Table 1.2):

Chemical change	Physical change
A new substance (or substances) produced	Same substance(s) before and after the change
(Usually) involves a large energy change	(Usually) does not involve a large energy change
Irreversible	Reversible
Involves breaking of bonds	Does not involve breaking of bonds

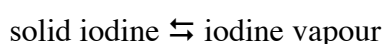
Table 1.2: Some criteria used to distinguish chemical and physical change

The first criterion listed, is the most fundamental, and is generally clear cut as long as the substances present before and after the change are known. If a new substance has been produced, it will almost certainly have different melting and boiling temperatures than the original substance. The others are much more dubious. Some chemical changes involve a great deal of energy being released: such as the example above of burning magnesium in air; or even require a considerable energy input, such as the example of electrolysis of water. However, other reactions may not obviously involve large energy transfers, for example when the enthalpy and entropy changes (see Chapter 4) more or less cancel. The rusting of iron is a chemical reaction, but usually occurs so slowly that it is not apparent whether the process involves much energy transfer.

It is important to be careful with language when discussing energy changes: energy is always conserved, and if we talk about a reaction ‘producing’ energy, that might seem to suggest the energy has been created rather than *converted* from another form. The physics teachers in the school may have guidelines on the preferred way of talking about energy ‘transfer’ or ‘conversion’, and it may help students if a common approach is used across the sciences.

Generally speaking, physical changes are more readily reversible than chemical changes. However, again this is not a very definitive criterion. The idea that chemical reactions tend to either 'go' or not is a useful approximation, but there are many examples of reactions that can be readily reversed (see Chapter 4). In principle, *all* reactions involve equilibria of forward and reverse reactions, and can be reversed by changing the conditions sufficiently. When hydrogen and oxygen are exploded it takes a pedant to claim that there is *also* a process of water molecules being converted into oxygen and hydrogen molecules as the reaction proceeds, which means the reaction will continue for ever. *Technically* such a claim may be true: but for all practical purposes the explosion reflects a reaction that very quickly goes to completion.

One technique that can be used to separate iodine from sand is to gently warm the mixture in an evaporating basin, over which is placed an upturned beaker or funnel. The iodine will sublime – turn to vapour, before re-condensing on the cold glass, separated from the sand. The same technique may be used if ammonium chloride is mixed with the sand. In both cases the separation is achieved because sand (which has a high melting temperature) is mixed with another substance in the solid state that is readily changed into a vapour by warming, and then readily recovered as a solid sample when the vapour is in contact with a colder surface. There are then reversible changes involved in both cases:



In the first case, the process involves only changes of state: evaporation and condensation - collectively called sublimation. However the second case involves one substance (a salt) changing to two other substances. To a student seeing these changes demonstrated, there would be little basis to infer one is (usually considered as, see Chapter 3) a chemical change, but not the other.

(It is worth noting that iodine, ammonia and hydrogen chloride are all rather unpleasant substances – and if this demonstration is attempted a proper risk assessment is needed.)

The final criterion in Table 1.2 concerns whether bonds are broken and made during a change, and this can only be meaningful for students once they have learnt about particle models of the submicroscopic structure of matter (see Chapter 2). In a chemical change, there will be the breaking of bonds that hold together the reactants, and the formation of new bonds in the products. However, we have to be careful here what we mean by 'bond' (this is discussed further in Chapter 3).

When ice melts and water boils there are 'intermolecular' forces between molecules disrupted, and this includes the breaking of hydrogen 'bonds'. However, when people talk about bond breaking in the context of chemical and physical changes they tend to mean strong chemical bonds such as covalent, ionic and metallic bonds (see Chapter 3). Yet even this is not clear-cut. When metals evaporate or are boiled, metallic bonds are broken, although the vapour is not normally considered a different substance.

When elements such as carbon and phosphorus undergo phase changes relating to allotropy, there is breaking, and forming of bonds, which might suggest these changes are chemical, and that the different forms of the same elements should be considered different substances. As suggested above, the status of different allotropes of the same element introduces a complication that is probably best avoided when first introducing the key ideas of chemistry.

Allotropes occur where different structures are more stable under different conditions of temperature and pressure. So in conditions deep in the earth the most stable form of carbon is as diamond. However at the earth's surface, diamond is less stable than graphite. Diamonds are 'meta-stable', which means that although they theoretically change to graphite after being brought to the surface, this is a very slow process. Chemistry tells us that diamonds are not 'for ever' - but this need not worry us mortals.

A particularly tricky case occurs when we dissolve materials to form solutions, especially of materials with ionic bonding (see Chapter 3). Dissolving tends to involve small energy changes, and to be readily reversible, and is generally considered a physical change. However, to dissolve an ionic compound such as sodium chloride (table salt) the strong ionic bonds between the sodium and chloride ions have to be overcome (and new bonds form between the ions and solvent

molecules). This would seem to suggest dissolving can be a chemical change according to the criterion of bond breaking and formation (Table 1.2).

Teaching about chemical and physical change

It appears the distinction between chemical and physical changes is a rather messy one, with no clear criteria to help students understand the difference. Some chemistry teachers avoid the distinction, and consider that it is not useful. However, some school curricula and examination specifications do set this topic, as part of the target knowledge students should learn about. It does not make sense to teach chemical and physical changes as any kind of absolute distinction, as this would seem to be unsupportable in terms of the chemistry, and will ultimately involve pointless rote learning and/or be a source of frustration for students. However, the idea that a chemical change involves changes in substances *is* a key idea; which should be introduced early in school chemistry, and reiterated in suitable contexts as students develop sufficient background knowledge to increasingly understand the principle.

Talking about changes of state as being physical changes, because despite apparent material differences, the same substance is present after the change, is also a central teaching point (see Chapter 2). The distinction is worth emphasising in these terms whenever suitable contexts are met. Ideas about energy changes and reversibility can be introduced (especially if specified in the curriculum) but should be presented as ‘rules-of-thumb’, that is as heuristics that are often useful, but which can mislead us. Similarly, when students have learnt about different types of chemical bond, the extent to which it is helpful to think about bond breaking/formation as being indicative of chemical change can be explored. With many students this might best be limited to another useful rule-of-thumb, that breaking and making of strong chemical bonds is often associated with chemical changes, but can sometimes occur without a chemical change.

For older students ready for some challenge, this topic can form a basis of a useful discussion task that can be set up as group-work. However, the focus should be *the extent* to which the criteria suggested in Table 1.2 are useful in making a distinction. The task would not be about coming to the right answer, but rather be based on using

evidence (from their knowledge of chemistry, or resource material) to argue a case for their position. In this way a rather unsatisfactory topic becomes the basis for practicing scientific argumentation, reviewing knowledge of specific chemistry, and learning something about the nature of science. (Table 1.2 can be considered to present a *model* of changes studied in chemistry, which - like all models - has limitations.)

1.5 Stoichiometry

A key pattern in chemical reactions, alluded to above, concerns the constant ratios found in most chemical reactions. (Most, because sometimes the same reactants may give different products: iron and oxygen, for example, can form FeO, Fe₂O₃, or Fe₃O₄, or a mixture depending on conditions) So, in our example of magnesium and oxygen reaction considered above:

3g of magnesium will completely react with 2g of oxygen to produce 5g of magnesium oxide
6g of magnesium will completely react with 4g of oxygen to produce 10g of magnesium oxide
9g of magnesium will completely react with 6g of oxygen to produce 15g of magnesium oxide
12g of magnesium will completely react with 8g of oxygen to produce 20g of magnesium oxide
24g of magnesium will completely react with 16g of oxygen to produce 40g of magnesium oxide
60g of magnesium will completely react with 40g of oxygen to produce 100g of magnesium oxide
600g of magnesium will completely react with 400g of oxygen to produce 1kg of magnesium oxide
3kg of magnesium will completely react with 2kg of oxygen to produce 5kg of magnesium oxide
etc, or in general
1.5X of magnesium will completely react with X of oxygen to produce 2.5X of magnesium oxide

Although the ratios found in different reactions are constant, they are not always this simple.

1g of hydrogen will react with 8g of oxygen to form 9g of water
23g of sodium will react with 35.5g of chlorine to give 58.5g of sodium chloride
100g calcium carbonate will decompose on heating to give 56g calcium oxide and 44g of carbon dioxide
etc.

The existence of such ratios can be explained by the models chemists use of the structure of matter at the submicroscopic level, and indeed is part of the reason chemists often initially adopted such ideas in an instrumental way: that is, as ideas

that worked as useful tools for thinking about chemistry, but which might not reflect an underlying reality. Over many years these ideas were found to provide a central set of models that could provide a central unifying framework for making sense of chemistry. Today most chemists consider molecules, atoms, protons, electrons and so forth to be real objects: but it is important to remember these ideas are a set of theoretical models, even if a very useful and successful one. I am not suggesting these entities do not exist, but our scientific models of these entities are subtle and still being developed, and the mental models that most of us have of them are at best partial and approximate versions of the best descriptions science can currently offer.

It is certainly very important to teach these ideas as theoretical, because although the models are successful and central to modern chemistry, it is not helpful if students think our models of atoms and molecules are precise realistic descriptions. Certainly the models introduced at secondary level fall somewhat short of this. As just one example, the notion that atoms contain ‘shells’ of electrons should not be taken to imply either that there is any kind of physical shell which contains the electrons (as some students assume), nor that the electrons in a shell can always be considered as equivalent. Students who select chemistry as a subject for further study will soon run into problems if they develop fixed ideas along these lines. It is much better to teach that atoms often behave *as though* they have electrons arranged in shells, but to warn students that scientists have found this is a simplification. That provides students with a more authentic understanding, avoids over-commitment to the model that might impede more advanced learning, and better reflects the nature of chemistry as a science.

Teaching basic ‘particle’ theory, will be discussed in the next chapter (Chapter 2). The key ideas that will be needed to see how this explains stoichiometry include:

1. Matter is quantised: that is, at a submicroscopic level matter consists of myriad discrete bits. We often refer to these bits as ‘particles’ although this is an analogy with familiar bits of matter like salt grains or specks of dust. The quanta of matter are at a MUCH smaller scale, and also have strange properties that are not like familiar particles – sometimes two of them can be in the same space for example. If we use the term ‘particles’ with students, we need to make sure we are clear about this

comparison. Students have sometimes been found to be confused about whether salt and sugar grains are examples of these ‘particles’. I will refer to these quanta of matter as ‘quantics’ to emphasise the difference.

2. There are a small number of basic types of quantics of interest in chemistry. The latter qualification is useful because physicists will talk about a large zoo of different particles – some of which are highly unstable under normal conditions and can only be produced in very specialised (and expensive) high energy colliders. In terms of a model which is useful for teaching secondary chemistry, it is usually enough to know about protons, electrons and neutrons.

3. Protons, neutrons and electrons are usually found clumped together (except at very high temperatures), and it is these clumps that are often the level of ‘quantic’ most useful for discussing what is going on in chemistry. Protons and neutrons are bound together in nuclei (plural of nucleus) by what is known as the strong nuclear force, and electrons are attracted to nuclei because of their opposite electrical charges – electrons are negatively charged, and protons in the nuclei are positively charged. The clumps that form are usually neutral (molecules) or nearly neutral (ions) – simply because forces exist between oppositely charged quantics, attracting them together. Under most important conditions, ions tend to be found either in neutral lattices or surrounded by other quantics (so in aqueous solution, ions are surrounded by a sheath of water molecules that are attracted to the ion).

4. The electrons do not get attracted right into the nuclei (luckily, as when that happens you get a neutron star, where no chemical substances or normal materials can exist), but are associated with one or more nuclei according to some rather complex rules.

The following table (Table 1.3) outlines the key quantics we commonly talk about in explaining chemistry. It is worth remembering that quantics are not like familiar particles: they are more *like a* hybrid of a particle and a wave. The entries in the table are organised in terms of ‘clumpiness’. So protons, neutrons, collectively nucleons, and electrons can all be considered as single quanta of matter in chemistry. The nucleus is a clump of nucleons. The atomic core is a clump containing a single nucleus, surrounded by electrons, and is largest type of clump that is generally

unchanged in chemical processes. Molecules, atoms and ions are larger clumps which are modified in chemical processes, and which tend to be characteristic of particular substances.

Quanticle	Description	Extension notes
Electron	Considered a fundamental particle. Has a negative charge (-1). It is much less massive than nucleons (see below).	The electrical charge is about $1.6 \times 10^{-19} \text{C}$ in S.I. units. Electrons have inherent angular momentum ('spin'). The mass of an electron is about $9.1 \times 10^{-31} \text{kg}$. This is so much less than the mass of a proton or neutron it is usually negligible.
Nucleon	A collective term for the particles found in nuclei, i.e. protons and neutrons	Made up of three quarks. Mass is $1.7 \times 10^{-27} \text{kg}$ – often called one atomic mass unit
Proton	A positively charged entity that is attracted to other nucleons by the strong nuclear force	Made up of three quarks. Mass is $1.7 \times 10^{-27} \text{kg}$ - often called one atomic mass unit
Neutron	A neutral entity that is attracted to other nucleons by the strong nuclear force	Made up of three quarks. Mass is $1.7 \times 10^{-27} \text{kg}$ - often called one atomic mass unit
Nucleus	The clump of protons and neutrons at the centre of an atom	Nuclei are unchanged by chemical processes, although some are unstable and undergo radioactive decay
Atomic core	A nucleus, and any 'shells' of electrons that can be considered to be fully associated with that nucleus. (That is electrons <i>not</i> in the valence or outermost shell.)	In some countries this is called a kernel. In most chemical changes, the atomic core remains unchanged (whereas there are changes in the arrangements of outer electrons).
Molecule	A neutral entity comprising one or more atomic cores and an outer layer of electrons that electronically cancels the nuclear charge.	Note – in common use, the term molecule is sometimes reserved for species with two or more atomic cores: not 'monatomic molecules' such as He, Ne etc.
Atom	A neutral species with one nucleus – that is it has the same number of electrons as protons.	Most atoms have outer (valence) electron shells that are not symmetrical, and atoms are rarely found under normal conditions. The exceptions are the inert gases.
Ion	A charged entity with one or more atomic cores	Ions that are commonly found have particular valence shell electron arrangements
Simple ion	An ion with a single atomic core surrounded by a shell of electrons	Simple ions differ from atoms because they have either too many or too few electrons to cancel the nuclear charge.
Molecular ion	An ion with several atomic cores surrounded by an outer 'layer' of electrons	Molecular ions differ from molecules because they have either too many or too few electrons to cancel the nuclear charges.

**Table 1.3: Quanticles – Theoretical 'wave-particle' objects used
by chemists to describe and explain the submicroscopic
structure of matter**

Stoichiometry is largely explained in terms of the electronic configurations of the atoms of the elements, as stable species are those where the atomic cores are well-shielded from other species by the outer 'layer' of electrons. This tends to happen in ions, molecules or sometimes atoms, with particular patterns of valence electrons (see Chapter 3). This leads to atomic cores combining in fixed ratios. The evenness (symmetry) of the pattern of electron density is more important than the neutrality of the species, so ions such as Na^+ and O^{2-} are found in common materials, whereas the atoms Na and O are not. However, it is important to realise that the stability of ions is only possible because they are usually found in neutral lattices, or surrounded by solvent molecules: isolated ions are usually less stable than the corresponding atom.

So, for example, magnesia has the formula MgO because magnesium forms an ion, Mg^{2+} , and oxygen an ion, O^{2-} , both of which have a symmetrical pattern of charge, and which can be stabilised by being formed into a an MgO lattice (a great many Mg^{2+} alternating with just as many O^{2-}) which is neutral overall. (Although other ions can be formed, e.g. Mg^{3+} , O^- , etc, these structures are too unstable to be readily stabilised.) The reason why magnesium and oxygen most commonly form these particular ions (as well as why, for example, a molecule of ammonia has three hydrogen atomic cores and only one nitrogen atomic core) can be understood in terms of their electronic configurations. This is explained further in Chapter 3.

Models of the structure of the atom that scientists find useful vary considerably in complexity, but in introductory chemistry it is useful to think that the electrons around a nucleus are arranged in shells. The elements in successive groups across a period of the periodic table reflect increasing numbers of electrons in a shell, and the breaks for a new period reflect the starting of a new shell. (This is, alas, a simplification, as only the first two shells fill completely before a new shell is begun. So, in the third shell, only eight of the maximum of eighteen electrons are in place before the fourth shell is used, in the first element of period 4, potassium.) Versions of the periodic table that represent the electronic structures of the atoms of the elements (see figure 1.15) can be useful in appreciating the stoichiometry of chemical compounds

Figure 1.15: A section of the periodic table in terms of electronic structure of atoms

[Use figure 7.7 (p.229) of the first edition here.]

1.6 How much stuff do we have? The mole concept

The phenomenon of stoichiometry - that chemical reactions involve precise mass ratios of reactants, leading to precise mass ratios of products - is understood in terms of the models of the substances being comprised of quanta such as molecules and ions at the submicroscopic level. The ratios themselves relate to how different substances can be understood in terms of the composition of those substances at the submicroscopic level.

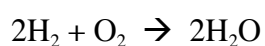
So we have just seen that magnesia has the formula MgO because it comprises of equal numbers of Mg²⁺ and O²⁻ ions. However, although the ion ratio is 1:1, this does not mean that equal masses of magnesium and oxygen react: rather we have seen that 3g of magnesium will completely react with 2g. This is because the nucleus of a magnesium ion is more massive than the nucleus of an oxygen ion. Each ion of magnesium, element number 12, has a nucleus containing 24 nucleons (12 protons and 12 neutrons), and each ion of oxygen, element number 8, has a nucleus containing 16 nucleons (8 protons and 8 neutrons). The 1:1 ratio of ions in this case relates to mass ratios of 24:16 (which happens to simplify nicely to 3:2 in this particular case). Actually, because of the presence of isotopes, a small proportion of these ions will have different numbers of neutrons in their nuclei, and so the precise values are not quite so neat!

The same principles apply to other examples, although the numbers are not always so convenient. So calcium fluoride has the formula of CaF₂, as the symmetrical ion of F is F⁻, and that of calcium is Ca²⁺ (see Figure 1.15). Therefore the crystal lattice of CaF₂ contains twice as many fluoride ions as calcium ions (as it will be neutral overall). Calcium is element 20, and the ion (ignoring isotopes) has mass number 40 (20 protons, 20 neutrons); whereas for fluorine the ion (again ignoring isotopes) has mass number 19 (9 protons, 10 neutrons), so the mass ratio for reacting calcium and fluorine is 40:(2×19)=38, so 40g of calcium will react with slightly less,

38g, of fluorine - and 20g of calcium will react with 19g of fluorine; 10g of calcium will react with 9.5g of fluorine, etc.

Using mass ratios allows us to scale up or down the amounts of reacting materials as much as we like (40 tonne of calcium will react with 38 tonne of fluorine; 40 μg of calcium will react with 38 μg of fluorine; etc). However, it is sometimes useful to have a standard way of talking about how material we are using, that allows us to shift directly from talking about individual ions or molecules to talking about laboratory scale amounts of material. For this chemists use a common scaling factor called the mole.

Consider the example of the reaction between hydrogen and oxygen: e.g. how two molecules of hydrogen will react with one molecule of oxygen:



The reacting mass ratio for hydrogen and oxygen therefore depends on two factors: the relative molecular masses of oxygen and hydrogen, and the ratio of molecules that react together.

As the mass of electrons are so small compared with the masses of protons and neutrons, and as neutrons are only very marginally heavier than protons, the reacting masses of substance follows very closely from a consideration of where the nucleons (neutrons and protons) are in the reacting substances. Consider the case of hydrogen and oxygen reacting (see figure 1.16).

Substance	hydrogen	oxygen	water
Molecules involved in reaction	2H_2	O_2	$2\text{H}_2\text{O}$
Nuclear composition	Two molecules of hydrogen, each with two nuclei containing one nucleon each (2×2)	One molecule, with two nuclei, each containing 16 nucleons (2×16)	Two molecules, each with three nuclei, containing 1, 1 and 16 nucleons $(2 \times \{1+1+16\})$
Total number of nucleons	4 nucleons	32 nucleons	36 nucleons

Figure 1.16: How reacting mass ratios depend on where the nucleons are

Literally, this tells us that 3.64×10^{-30} kg of hydrogen (two molecules) reacts with 29.12×10^{-30} kg of oxygen (one molecule) to produce 32.76×10^{-30} kg of water (two molecules). However, because mass is quantized, this also tells us that any reaction of hydrogen and oxygen will involve multiples of these masses as 2000 molecules of hydrogen will react with 1000 molecules of oxygen; 2 000 000 molecules of hydrogen will react with 1 000 000 molecules of oxygen; 2×10^{24} molecules of hydrogen will react with 1×10^{24} molecules of oxygen, and so forth.

The chemist's dozen

This is the context in which the idea of the mole, as a measure of the 'amount of substance' is used in chemistry. A lot of students find the idea of the mole difficult, and this is presumably due to the combination of two factors: the abstract nature of the concept, and the expectation that students will apply maths in a chemical context.

The important idea to get across to students is

(a) that we can explain stoichiometry (constant reacting ratios) in terms of the masses of the particular molecules and ions involved in the reactions concerned, or the 'relative' molecular masses if we allow for different isotopes – however,

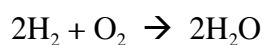
(b) that is not very practical in the laboratory, as chemists operate with samples that contain billions and billions of molecules, and need to measure out samples in a way they can easily manage (not by counting billions of molecules that are too small to be seen!)

So in most reactions, a useful common way of measuring reactants is in terms of mass: how many grammes, or kilogrammes, of substance is being reacted. At the bench level, the gramme is a more suitable starting point, so the mole is based on masses in grammes that reflect the number of nucleons present in a sample. So for hydrogen (two nucleons per molecule), one mole would be 2g; and one mole of oxygen (32 nucleons per molecule) is 32g. As two molecules of hydrogen react with each molecule of oxygen, two moles of hydrogen (4g) are required to react with one mole of oxygen (32g). Thus the reacting masses in grammes reflect the nucleon ratios in figure 1.13 above. So the mole can be understood as a multiplier like a dozen (12), or a score (20), or a gross (144), only a lot bigger. Indeed, as each nucleon only weighs 9.1×10^{-28} g, it turns out that a mole is about 6.0×10^{23} – a bit bigger than a standard dozen!

Substance	hydrogen	oxygen	water
Molecular formula	H ₂	O ₂	H ₂ O
Relative molecular mass	2	32	18
Stoichiometric ratio	2 H ₂	(1) O ₂	2 H ₂ O
Reacting mass ratio	2·2 = 4	32	2·18
Reacting masses (for reacting one mole of oxygen)	4g	32g	36g
Reacting masses (to generate one mole of water)	2g	16g	18g

Figure 1.17: How reacting mass ratios depend on where the nucleons are

The strength of describing reactions in terms of moles (abbreviated as mol) is that it can be scaled up or down by any amount:



2 mol : 1 mol : 2 mol

2·2g=4g : 32g : 2·18g=36g

10 mol : 5 mol : 10 mols

10·2g=20g : 5·32g=160g : 10·18g = 180g

0.1 mol : 0.05 mol : 0.1mol

0.1·2g=0.2g : 0.05·32g=1.6g : 0.1·18g=1.8g

etc.

We could carry out similar analysis in other examples, including for the formation of magnesium oxide or calcium fluoride, discussed above. An important teaching point is to note that we used the term ‘relative molecular mass’ as a general term, even when molecules are not involved. So magnesium oxide, a lattice containing ions, is said to have a relative atomic mass of 32.

Students seem to vary according to how they find it best to work out mole calculations (perhaps related to different ‘learning styles’ or ‘thinking styles’). Some prefer to use verbal arguments, some doing the algebra, and some using graphical schematics of various kinds. In general, it is found that learning of scientific ideas is supported when multiple forms of representation are used. It would seem sensible to model different approaches, and allow students to find something they are comfortable with. This could be introduced in a dialogic form by setting a task that students are asked to work on in pairs or small groups, and then asking each group to explain their approach to the class and inviting comments on the strengths and weaknesses of each approach. You can add any key approaches not suggested by the class. Given that many students struggle when asked to work using a particular formalism or set of rules preferred by the teacher this could be a good use of time: it will show students that there are different ways of thinking about mole problems, and indicate that several different approaches can be used to get the right answer (they are all equivalent at a fundamental level of course!) This can encourage them to think about the logic of what they are doing, rather than trying to learn an algorithm that will fail as soon as they have a slightly different form of task to complete. This will encourage them to value understanding over rote learning of rules-of-thumb. This approach also reinforces how in chemistry there are often alternative ways of representing the same information – for research suggests that students tend to make better progress when they have a repertoire of alternative representations they can draw upon to do their chemical thinking. Finally, if this seems a rather brave way of approaching using maths in chemistry classes, it might be worth talking to the maths staff. In many schools this general approach of getting small groups to work on maths problems and then share and discuss alternative methods is a core feature of maths lessons.

Range of moles problems

At secondary level, students are usually only asked to address moles questions of limited complexity. However, for students that progress in the subject, the demand of problems can get quite high.

As well as calculating the masses of reactants or products as in the example above (where the complexities of the mass ratios can vary according to the reaction, of course), students can be asked to solve problems with gases (where the volume of a gas at particular temperature and pressure depends upon the number of moles present, i.e. Avogadro's principle) and solutions (where the amount of solution is measured by volume, and a conversion according to molarity, the number of moles per unit volume of solution, has to be made).

As with most of the topics discussed in this chapter, the mole is a fundamental idea in chemistry, which is best introduced with fairly simple examples, then revisited over an extended period of time. In terms of introducing the topic, it is worth noting that the sheer range of possible examples (any chemical equation can be used as the basis; any feasible value for reacting masses can be used) means that such a topic lends itself to ready differentiation for students. Where some students may struggle with examples involving simple ratios and masses chosen to give simple whole number answers, others may soon be ready for more demanding examples with complex mole ratios, and arbitrary choices of reacting masses to be calculated to several decimal places.

Coda

This initial chapter has introduced a range of key concepts, most of which are abstract and are known to be challenging for many learners. However, these topics are central to chemistry, and essential for a modern understanding of the subject which goes beyond simple description of substances and their reactions.

Individually the notions of elements and compounds, of theoretical models of structure at submicroscopic level, of chemical change, of chemical equations and stoichiometry, and of the mole, are all likely to be found difficult for many students.

Despite this, however, as these ideas become familiar they also become mutually supporting as they build to give a coherent way of thinking about chemistry that is at the core of the wide range of chemical explanations that has made chemistry the successful science it is: developing new materials to make life healthier, safer, more comfortable, and more entertaining. Teaching these ideas is not the work of a small number of discrete lessons, but should rather be the basis of teaching over the secondary years – being regularly reinforced with new examples in new contexts. In this way what are initially odd and abstract ideas to be learnt and remembered can become a set of familiar thinking tools that link together to turn chemistry from being an amazing but mysterious subject of colours, smells and bangs: into a coherent way of making sense of the material world.

Further reading and resources

The Wellcome Foundation has a collection of images that can be used in teaching under a Creative Commons license. For example, the collection includes drawings of Lavoisier at work, and of Priestly's chemical apparatus.

<http://images.wellcome.ac.uk/indexplus/page/Home.html>

CLEAPSS - Support for practical work, and in particular health and safety information for school science. CLEAPSS is the source for such useful resources as the *Secondary Science Laboratory Handbook* and *Secondary Science Hazcards* (providing Safety information and model risk assessments for handling chemicals)

<http://www.cleapss.org.uk/>

A number of useful tools about the elements and the periodic table can be found on the Royal Society of Chemistry's site: e.g. 'visual elements' and the 'periodic table of data'.

<http://www.rsc.org/Education/Teachers/Resources/OnlineResourcesHome.asp>

The University of Nottingham's 'periodic table of videos', a series of short films discussing different elements, may be found at

<http://www.periodicvideos.com/about.htm>