

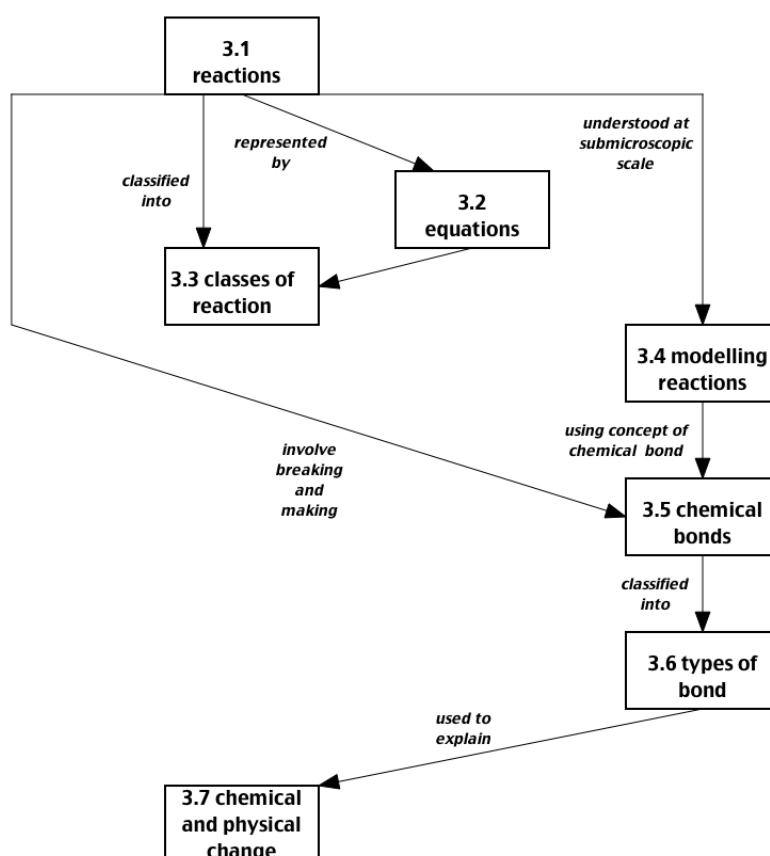
# Teaching Secondary Chemistry

## Chapter 3

### Introducing chemical change and chemical bonding

Keith S Taber

#### Topic Schematic



This is the author's manuscript version.

The version of record is:  
Taber, K. S. (2012). Introducing chemical change.  
In K. S. Taber (Ed.), Teaching Secondary  
Chemistry (2nd ed., pp. 75-101).  
London: Hodder Education.

## ***Previous knowledge and experience***

Students will come to secondary school with experience of chemical reactions such as combustion (gas cookers, bonfires etc), and aware that cooking brings about changes in food stuffs (unfortunately rather complex examples), as well as knowing that some household ‘chemicals’ such as bleach are used to ‘kill germs’. Unfortunately most experience in everyday life does not provide a clear basis for characterising chemical changes, with many household ‘chemicals’ (which are usually mixtures of different substances, see Chapter 1), such as washing powders, largely working without chemical change (but rather facilitating separation techniques). Moreover, not only are chemical changes not a clear distinct category of phenomena, but students often tend to think of changes in terms of an active agent (e.g. the bleach) which acts upon a substrate, which is not helpful for thinking about the reactions between two substances.

Chemical bonding is a theoretical idea, that is part of the way matter is modelled by chemists at submicroscopic scales, (as composed of extremely tiny ‘quanticles’, see Chapter 1) and so not likely to be familiar from outside the classroom. However, if particle models are introduced in the manner recommended in this handbook (see Chapter 2), students will have been introduced to the idea that the particles of which substances are made have ‘holding power’ of varying strengths, which tends to cause them to clump together unless they have sufficient energy of movement to overcome this.

## ***Choosing a route***

The comments made earlier in the book (in Chapter 1) about teaching chemistry in an iterative fashion apply here as well (and indeed there is some inevitable overlap in developing basic chemical ideas across Chapters 1-3). The topics included in this chapter have been arranged in an apparently logical order, to build-up an understanding of the basic ideas relating to the key concept of chemical reaction. However, the teaching needs to be iterative, and to follow a spiral curriculum: learning about each of the key concepts of this Chapter will be supported by, and will support, learning about the others.

Moreover, even if students appear to have a good grasp of these ideas when they are first introduced, research suggests that abstract concepts such as this will over time become muddled unless there is careful regular reinforcement. Luckily, the ideas met in this Chapter (as with the other Chapters in this section of the book) have applications across the teaching of chemistry. It is important, therefore, to explicitly refer to these ideas in teaching other topics, to regularly review understanding of the key ideas. Indeed, this is necessary to ensure that students are making the links that chemists and science teachers take for granted. Again, research suggests that what seems obvious to teachers (because we have previously stressed how it always applies in chemistry; because the technical language we use should clearly imply it) is readily missed by many students. So when teaching about topics such as acids, geochemistry, chemical analysis and so forth, it is useful to always check that students are thinking about the new teaching in terms of the basic concepts of the subject.

The route set out here reflects a very common aspect of teaching and learning chemistry – the shift between (i) macroscopic (bench) phenomena; (ii) their formal representation in conventional symbolic language (formulae, chemical equations); and (iii) the explanatory models that chemists use that are based on theoretical ideas about the structure of matter at a sub-microscopic scale (molecules, ions, etc). The chapter schematic has been set out to illustrate the main relationships between these three levels of communicating about chemistry (the topics are organised into three columns). Whatever route is taken through this material, it is important to signpost the shifts between these three levels, to help students begin to appreciate how chemistry explains the material world by working at these three distinct ‘levels’.

### ***3.1 Chemical reactions***

The distinction between physical and chemical change was discussed in Chapter 1. There it was suggested that although this is not a clear-cut distinction, it can be a useful simplification. In chemical terms, the distinction concerns whether there is a change of substance: in a physical change the same substance changes its form or state, but in a chemical change there is/are different substances present after the change.

In chemistry we refer to chemical changes as ‘reactions’ or ‘chemical reactions’. Like many terms used in a technical sense in the sciences, ‘reaction’ can have unhelpful associations for students. The term react can imply a response to something, and research suggests that for many students a chemical reaction is understood as one chemical in some sense provoking a reaction in another. That is, one chemical is seen as being the active substance, bringing about change, whilst the other is more a victim of chemical intimidation! For example, when acids react with other substances, students may assume that it is the acid that is actively bringing about the reaction in the other substance. Students may also see the reaction as being *to* heating or stirring, or even as a reaction to the chemist adding a reactant. It may be useful to adopt a class motto along the lines ‘reacting with, not reacting to’.

Colleagues teaching physics also use the term ‘reaction’ in relation to forces. Whenever there is an interaction between two bodies (a person standing on the earth for example) there is equal magnitude force acting on both bodies. This has traditionally been called ‘action’ and ‘reaction’, but any implication that one body initiates the interaction and the other responds is completely wrong: the interaction is always mutual, without any time lag in the symmetry.

The tendency of students to see chemical reactions in a similar way, initiated by one partner, is perhaps not as problematic. After all, chemists often think and talk in terms of ‘attacking’ species in reaction mechanisms. However, there are some chemical reactions where this does cause a problem. Consider the following change, which occurs when copper carbonate powder is strongly heated:



This is an example of a decomposition reaction (considered further later in the chapter). This is a chemical change as the substance present at the start (copper carbonate, a green solid) is no longer present after the change. Instead two new substances have been produced: black copper oxide powder and invisible carbon dioxide gas. Copper carbonate is ‘reacting’, but it is not *reacting to* another chemical substance.

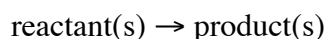
Thermal decomposition reflects how material tends to become less aggregated at higher temperatures (cf. Chapter 2): condensed matter (substances in the solid or liquid state) will vaporise if heated sufficiently; complex molecules will break down to simpler ones or fragments, and if we kept on heating we can atomise materials, ionise the atoms - and in theory sufficient heating would actually decompose nuclei and even the nucleons themselves. Chemists do not have to worry about these latter stages (as they only occur under extreme conditions, such as in the very early universe), but it is important to recognise that when we talk about the ‘stability’ of a substance or a chemical species, we should always bear in mind the conditions being referred to. At room temperature copper carbonate seems perfectly stable, but at a higher temperature (as reached by a Bunsen burner flame) it spontaneously decomposes.

Some students will therefore have difficulty considering thermal decomposition as a ‘proper’ chemical reaction, as the copper carbonate does not react with any other substance. Other students will not have this difficulty, but because they think of the powder ‘reacting to’ the heat. The things to watch out for here is if any students confuse substances and energy – that is, if they consider that heat can be considered as a substance. It is important that we emphasise that energy and substances are quite distinct, even though changes in the latter tend to involve changes (transformations) in the former.

As a teacher, it is useful to continually emphasise that the key criterion for a chemical reaction is *change of substance*, regardless of how many substances are involved before or after the change. So the thermal decomposition of copper carbonate is a chemical reaction because the products, copper oxide, and carbon dioxide, present after the change, are different substances to the copper carbonate that has reacted and no longer exists.

### 3.2 Chemical equations

Chemical equations are ubiquitous in chemistry (and have already been used in this Handbook), because they are such useful tools for communicating between scientists. A chemical equation summarises a particular chemical reaction: and such reactions are the central phenomena in chemistry. The general form of an equation is of the form:

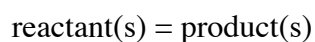


where the substances present before the reaction, the reactants, are signified (by the ‘→’ sign) to change into the substances present after the reaction, the products.

#### Key teaching points

*What exactly is equal in the equation?*

The term ‘equation’ might lead students to expect us to use an equal sign, something like:



but the normal practice is to use an arrow to show the direction of the change. The term equation reminds us that there is a sense in which the two sides of the ‘equation’ are equal. The total mass of products must be the same as the total mass of reactants (as we saw in Chapter 1). If we think in terms of the chemist’s models of the structure of matter at the submicroscopic scale, then we can understand that there is a conservation of the most basic entities: there are the same number of electrons, protons, and neutrons after the reaction as before (which in terms of this quantile model, can be understood as *the reason* that mass must be conserved). It is useful to reinforce this point for students: conservation of the ‘quantiles’ present, each having a fixed mass, implies conservation of mass in the reaction. The type of analysis presented in Chapter 1 (§1.6) may be worth revisiting at various times in the context of different examples of chemical reactions.

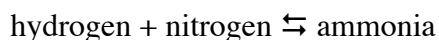
### *What about energy changes?*

As well as mass being conserved, so is energy. This is not something we tend to emphasise strongly in chemistry, but we should be aware that when our students study physics topics they will be taught that energy is never created or destroyed in any process. So although chemical reactions involve energy changes, sometimes very obvious and significant ones, the total energy of the system does not change.

Chemical systems are considered to have different amounts of chemical potential energy, and during reactions there may be a net conversion of chemical potential energy to (or sometimes from) other forms of energy – in particular kinetic energy of the new particles formed, reflected in the products often being at a higher temperature. This is usually only discussed in chemistry lessons near the end of the secondary years, but it is important that when we teach younger students we do not seem to be telling them something inconsistent with what they are learning in other science lessons. So when we talk about energy changes involved in chemical reactions we should be careful to refer to the energy transformed or transferred, and not to suggest that chemical reactions can ‘make’ or generate energy.

### *Directionality of a reaction*

We write chemical reactions as going from reactants to products, and this is a useful way of thinking that fits most reactions we come across (it is a good *model* for most reactions). However as students advance through the school, they will meet reactions that do not go ‘to completion’, such as in the industrially important manufacture of ammonia:



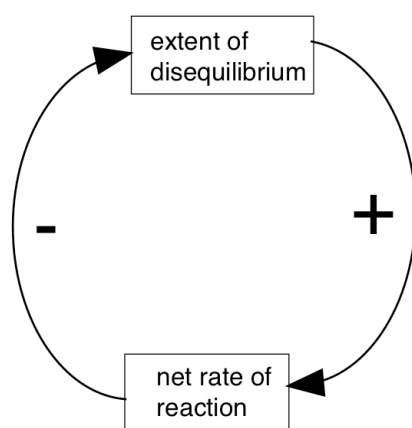
The decision to write the reaction as above, rather than as



is somewhat arbitrary in terms of chemical principles, but reflects *our use* of this reaction to make ammonia from nitrogen and hydrogen. We manipulate the conditions to produce ammonia, which is useful for making fertilisers etc, from the more readily available hydrogen and nitrogen.

There is a key teaching point here: all reactions are *to some extent* reversible. Even in a reaction such as exploding an oxygen and hydrogen mixture, it is possible theoretically to calculate some extent of reverse reaction (even if it would be of no practical importance at all in such an extreme case). More advanced students will study equilibria (see Chapter 4), from which perspective we consider that for any possible reaction, the balance between reactants and products found ‘after’ a reaction will depend upon the conditions of temperature, pressure etc. The word ‘after’ is placed in inverted commas here, as when reactions reach an equilibrium, there is still plenty going on at the submicroscopic level, but just no further net change (as forward and reverse reactions proceed at the same rate and cancel each other out).

It should also be noted that theoretically the approach to equilibrium will involve an ever decreasing rate of change, although often there is soon a point when there is no further observable change. We can understand this in terms of a simple feedback cycle (see figure 3.1). If we consider the driver for change to be the extent to which the reaction mixture is ‘out of equilibrium’, then the further from equilibrium the mixture, the greater the net rate of reaction (shown by a ‘+’ symbol) – which of course moves the reaction closer to equilibrium. However, that reduces the driver, so the rate of reaction slows, so the rate at which the reaction approaches equilibrium also slows.



**Figure 3.1: The approach to equilibrium is subject to a negative feedback cycle**

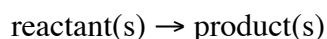


We do not commonly teach about feedback cycles in this way, but they can be a useful way to think about an abstract explanation with older students. The feedback cycle offers a useful tool for thinking about systems that links across the sciences. Homeostasis in the body depends upon such cycles. In physics there are a number of phenomena that can be described with the same very simple feedback structure as in Figure 3.1 (cooling of a hot object, radioactive decay, capacitor discharge, for example). And feedback cycles are very useful in thinking about aspects of earth and environmental sciences (for example the possibility of a positive feedback cycle when atmospheric warming leads to warmer seas in which the greenhouse gas carbon dioxide is less soluble).

In introducing ideas about reactions to younger students, we do not want to overburden them with such considerations as equilibria, as there are many examples of reactions where it is clearly sensible to designate reactants (what we start with) and products (what we end up with), and where for all practical purposes the reaction goes to completion, and can be readily observed to do so. However, in following the principle that we should avoid students developing ideas that will make progression in learning difficult, it is probably useful to make sure they are aware that some reactions never go to completion, and also that some reactions only occur very slowly under common conditions. More detailed advice on teaching these areas can be found in Chapter 4.

### **The chemical equation as a model**

This leads to the suggestion that the best way of thinking about the chemical equation



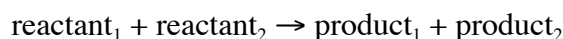
is as a *model*. Chemists and other scientists commonly used models which they know simplify and or over-generalise, where they are useful for thinking about aspects of nature. This equation is a very useful model because we can use it to describe any reactions. However, like all models it has its limitations. So for some reactions, those which are *more obviously* equilibria and where there will be some ‘reactant’ present no matter how long students observe the reaction mixture, the model has to be amended to reflect the equilibria. For the majority of reactions, although they are

technically equilibria, the model does a good job of describing the reaction for most purposes.

Presenting the idea of the chemical equation as a very useful model that can help us summarise what is going on in reactions, and which does a good enough job for most reactions, will both enable students to see the value of the formalism, and stop them making unfortunate generalisations about all reactions going to completion (which could then later act as misconceptions when they meet reactions that do not fit the model).

### Forms of reaction

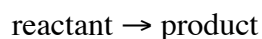
Various forms of reaction equation are possible, depending upon the number of different reactant and product substances, i.e.



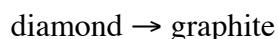
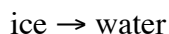
etc

Some classes will accept these different possibilities readily, and will not need much class time spent discussing this. For other students, this may seem quite complicated. A useful sorting and classifying activity for these learners would be to present them with a set of word equations for a range of reactions that they will come across during their course, and have them work in pairs or small groups to see which of these reactions fit into the different patterns. This will help students become familiar with the idea of using equations to represent chemical reactions. If you have a class with a wide range of abilities, you can set this task for some students, and ask others students (for whom this would be a trivial activity) to identify examples of reactions not on your list in the different categories from a text book (or the internet if readily accessible in your teaching room).

A good question to ask students is whether we could get chemical reactions of the form:



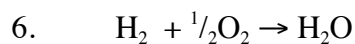
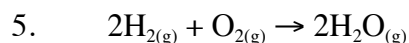
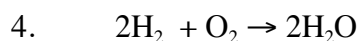
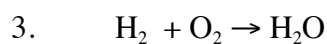
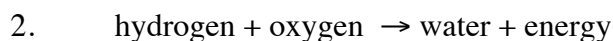
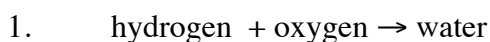
We can suggest *changes* that would fit this general pattern:



In these examples one *material* changes into another. However ice and water are different states of the *same substance*, and this example is a physical change (melting, see Chapter 1). Diamond and graphite are very *different materials*, but both forms of carbon - different allotropes. Again this is not usually considered a chemical change (although there is a change in the bonding within the structure when the carbon changes form), but this example does highlight how our basic chemical concepts run into difficulties in some situations (see Chapter 1).

### Word and formulae equations

Two most common types of chemical equation are those written in words using the names of the substances; and those representing the formulae of the substances involved (later students will meet representations involving structural formulae, see Chapter 8). Consider the following examples:



Each of these equations describes the reaction between hydrogen and oxygen to give water (although example 3, would normally be considered inadequate, as discussed below). Examples 1 and 2 are word equations, whereas examples 3-6 are formulae equations.

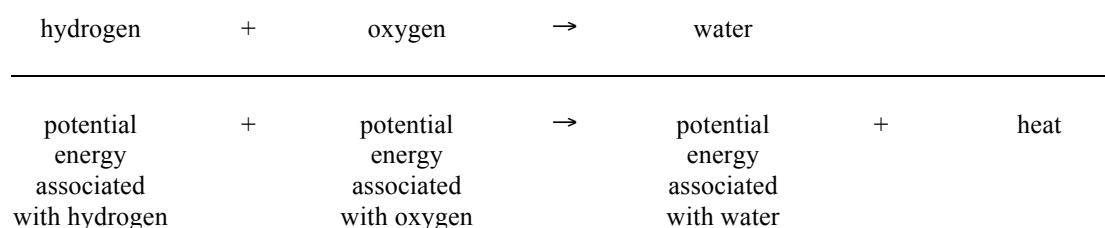
### *Word or formulae equations?*

It is sometimes thought that it is better to first introduce word equations with younger students, and then later progress to formulae equations. This can seem sensible, as formulae equations look more abstract and technical. However, formulae equations may have some advantages, even with younger students. For one thing, some common names of substances (water, ammonia for example) do not offer much clue to whether a substance is an element or a compound - and if a compound, which elements it is a compound of. Formulae equations are much more explicit here, even if looking more technical and unfamiliar to younger learners. Research has shown that when students are asked to complete simple word equations (describing types of reactions familiar from school science, and with just one word missing), many struggle to know how to proceed: and many of the correct answers are lucky guesses, or based on chemically very dubious logic. Formulae equations are much more explicit about the elements represented in the reaction, and allow students to readily check if everything on one side is present on the other. (As pointed out in Chapter 1, we need to be careful in how we talk about elements 'present in' compounds – see §1.3.) The assumption that word equations are easier is probably misguided, at least for some students.

The recommendation I would make here, is to introduce formulae equations as early as possible, and use them to *complement* word equations, so that younger students get used to seeing the same reaction represented in both forms. This looks like 'more' information for students to handle, but actually once they are used to this approach, the two formats are mutually reinforcing which will prove helpful to most students.

Example 2 above, refers to energy in the equation. Some textbooks do this to emphasise how some reactions release (note, not create) large quantities of energy.

Alternatively the term ‘heat’ might be used. We know that many students at lower secondary level do not distinguish well between materials and energy. For example, heat may be thought of a substance, as a kind of fluid (which is how it is sometimes modelled). Therefore examples such as this should be avoided, unless a clear formalism is used that students know indicates that energy is not a chemical substance (for example, always writing the energy term in the same contrasting colour). However, technically it does not make sense to include an energy term like this, as we should consider the energy separately. We might think of this in the following terms:



When we use a form of representation such as this, it is clearer that the heat generated is produced because of the difference in the levels of chemical potential energy in the reactants and products: energy is changing its form, but not being created.

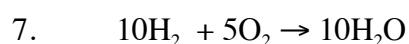
Some (mainly older) students may have heard that mass can be considered a form of energy for some purposes, and that changes in mass and energy are intimately related through Einstein’s famous equation  $E=mc^2$ . It is technically the case that in any chemical reaction where there is an energy change, there will also be a calculable change in the mass of the materials present (just as it is technically the case that warming a beaker of water changes its mass). However these effects are much too small to be of any significance at all in the chemistry laboratory, and for all practical purposes, mass and energy can be considered to be - separately - conserved in chemical changes. Any precocious student who suggests otherwise might be asked to do some research into the magnitude of the ‘mass defect’ that can be calculated for a typical reaction involving molar quantities (e.g. 2 mol of hydrogen reacting with 1 mol of oxygen).

Example 3 above is an unbalanced formulae equation. This shows the formulae of the substances involved in the reaction, but has not been ‘balanced’ to make sure the amount of each element is the same on both sides. That needs to be done to work out

the ratio of the reacting masses of the reactants (see Chapter 1). Although there may well be points in lessons when such (in)equations may be considered whilst developing ideas, it is strongly suggested that teachers should always then move on the balanced equation (here, example 4) when representing reactions with formulae in this way.

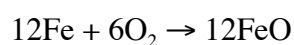
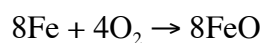
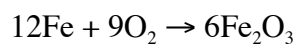
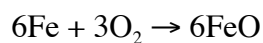
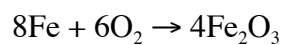
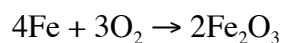
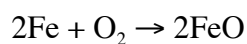
The difference between examples 4 and 5 is the level of detail provided. Example 5 offers additional information: the physical state in which substances are involved in the reaction. This is useful information when thinking about reactions we will observe in the laboratory, but often with younger students we will omit this detail when using equation to represent reactions, to keep the focus on the substances.

The difference between examples 4 and 6 is trivial for a chemist, as reaction equations show the mole ratios of substances involved, and multiplying throughout does not change this. So, for example, consider the following equation:



Equation 7 is not substantially different from equations 4 and 6 – all represent the same reaction, as do an infinite number of other possible versions! We would not usually use equation 7, as it is not the simplest way of writing the equation. Equation 4 is more common as it is the simplest way of writing the equation that only needs integers (cf. the  $\frac{1}{2}$  in equation 6). However, sometimes equation 6 is used (as two of the three substances involved are present as one unit).

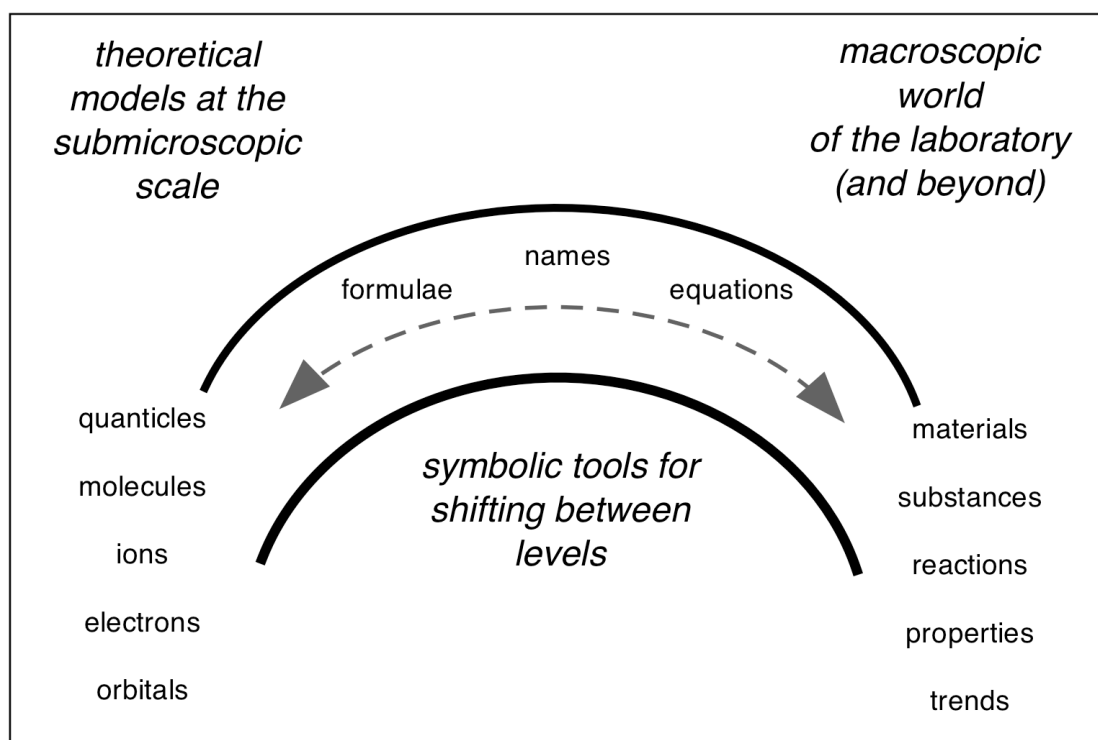
If some students find these ideas confusing, it might be worth providing them with variations (like, 4, 6 and 7 in this example) of equations for several reactions they will meet in their course, and asking them to work in groups to identify (a) which examples represent the same reaction, and (b) for each reaction, which representation is the simplest version. This should be a simple task for most learners, and can readily be tweaked to provide differentiated versions to challenge the most able students in a class. For example, distractors which have the wrong ratios could be included for higher attaining students, whereas only valid equations are included for less confident learners. The task becomes more difficult if it includes versions of similar but distinct reaction equations, e.g.



### **The equation as a bridge between molecules and moles**

One of the most important features of modern chemistry is that it is a science that provides explanations of the reactions seen in the world at observable scales (at the molar or macroscopic level) in terms of theoretical submicroscopic entities such as molecules and ions and electrons. As chemistry teachers we are constantly shifting back and forth between the macroscopic descriptions and submicroscopic models: this shifting takes place in our thinking, and is reflected in our classroom talk. To follow our arguments and explanations, our students need to be able to follow these shifts.

The symbolic language of chemistry acts as a bridge for us (see figure 3.2).



**Figure 3.2: Chemistry teachers use symbolic tools to shift student thinking between what they can see and the explanatory models used to understand observations**

When we speak words like ‘hydrogen’, ‘copper’, or when we write on a board ‘ $\text{H}_2\text{O}$ ’ or ‘ $\text{CuSO}_4$ ’ or ‘hydrogen + oxygen  $\rightarrow$  water’, we are of course using symbols. However, a special property of these symbols is that they are *usefully ambiguous*. They are ambiguous symbols because we could be referring to actual samples of substances at a scale we can manipulate in the classroom laboratory; or we could be referring to individual quanticles. ‘Hydrogen’ could mean the element in abstract, an actual sample of gas we have collected, or a single molecule. Indeed in the sentence ‘hydrogen only has one proton’ we would be referring to none of these, but rather a single atom of hydrogen.

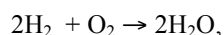
This is incredibly useful because it enables us as teachers to make shifts between the observations we want to explain and the explanations chemists develop in terms of theoretical models at the submicroscopic level. However, we may get so used to talking in this way, that we can easily forget that students may not always follow these shifts in thinking, and what we say may seem like some magician’s sleight-of-hand. So we might be talking about an actual reaction students have seen, and then we



summarise this by writing (or projecting) a chemical equation, where the terms represent the substances the students have observed in the reaction. Then we start talking about how, for example, the lone pairs on one species are attracted to the positive region on another: and now the same equation is referring to something rather different.

This tool of having labels that apply to both substances and molecules; and equations that refer to both bench scale reactions and reaction mechanisms at the level of individual molecules and ions, is incredibly powerful – but it also has great potential to confuse students if they do not spot when we have crossed the symbolic bridge (Figure 3.2) to shift between levels. As teachers, therefore, we need to be very explicit about using this tool, so that students not only realise when we make these shifts, but come to see the power of this tool to help their own thinking. I strongly suggest modelling the use of this bridging process in teacher talk:

“so we have seen the hydrogen combusts with a squeaky pop and we can write this as



which describes the chemical change. In the reaction the substances hydrogen and oxygen are changed into a new substance, water. The large numbers tell us we need twice as much hydrogen as oxygen in this reaction, but we must remember that is measured in moles.

The equation can also tell us why this is, because *it can also* summarise

what is happening to at the level of molecules. We see that two molecules of hydrogen are needed for each molecule of oxygen, to produce two molecules of water...”

### ***3.3 Common types of reaction***

Chemistry as a science makes use of a range of classification schemes. The periodic table (see Chapter 1) offers an excellent example of this, assigning elements to

'blocks', groups and periods. Reagents are characterised into types (acids, oxidising agents etc), and reactions can also be classified in similar ways. Given that there are an almost unlimited number of substances that can be formed in chemistry, and most undergo a range of reactions with various other substances, the sheer number of possible reactions is immense.

However, the ability to classify elements and compounds into groups according to similarity in properties (reaction behaviour) allows us to also classify reactions into useful groups. There are many specific named reactions in more advanced chemistry (such as the Wittig and Diels-Alder reactions), but even in introductory chemistry it is useful to set out some common types of reaction that students will meet in their studies. Every reaction will have some distinct, unique features, but a classification scheme provides a useful starting point for dealing with what would otherwise be an overload of information for the chemist (let alone the school student).

### **Common types of reaction**

The reaction types suggested here will link to many of the reactions met in school science or chemistry courses, and in each case we can write a *general* reaction equation that shows the types of substances involved in that type of reaction. In specific cases, the general labels for types of substance (e.g. metal, acid) are substituted by different examples. It should be noted though that in some general reactions there are common specific products (e.g. an acid reacting with a metal usually produces hydrogen, although the other product varies depending on the acid and the metal reacting).

### **Binary synthesis**

The formation of a compound from two elements:

General form	<b>element</b>	+	<b>element</b>	→	<b>compound</b>
e.g.	hydrogen	+	oxygen	→	water
	hydrogen	+	nitrogen	⇌	ammonia

## Decomposition

The breaking down of a compound into simpler products (sometimes elements, not always) on heating:

General form	<b>compound</b>	→	<b>element/compound</b>	+	<b>element/compound</b>
e.g.	copper carbonate	→	copper oxide	+	carbon dioxide
	ammonia	⇌	hydrogen	+	nitrogen

## Neutralisation (acid-alkali)

The term neutralisation is used to refer to acid-base reactions, and so can include reactions between acids and metal oxides and carbonates (which are basic), but in introductory chemistry usually refers to acid-alkali reactions. (See Chapter 5 for a discussion of acids, alkalis and bases).

General form	<b>acid</b>	+	<b>alkali</b>	→	<b>salt</b>	+	water
e.g.	hydrochloric acid	+	sodium hydroxide	→	sodium chloride	+	water
	nitric acid	+	potassium hydroxide	→	potassium nitrate	+	water

## Acid-metal

General form	<b>acid</b>	+	<b>metal</b>	→	<b>salt</b>	+	hydrogen
e.g.	hydrochloric acid	+	zinc	→	zinc chloride	+	hydrogen
	nitric acid	+	magnesium	→	magnesium nitrate	+	hydrogen

## Acid-carbonate

General form	<b>acid</b>	+	<b>carbonate</b>	→	<b>salt</b>	+	water	+	carbon dioxide
e.g.	hydrochloric acid	+	zinc carbonate	→	zinc chloride	+	water	+	carbon dioxide
	nitric acid	+	magnesium carbonate	→	magnesium nitrate	+	water	+	carbon dioxide

## Acid-metal oxide

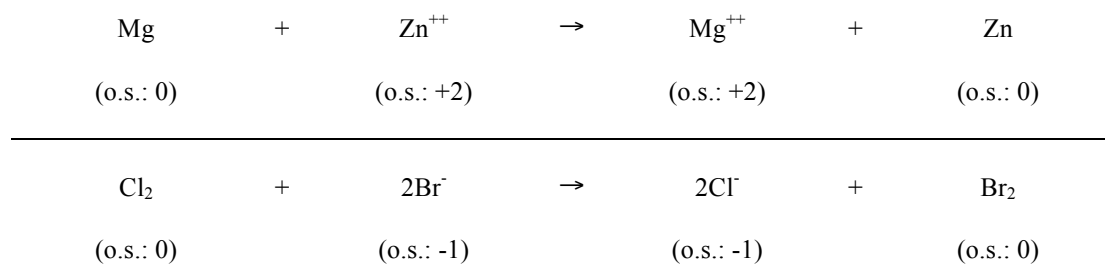
General form	<b>acid</b>	+	<b>oxide</b>	→	<b>salt</b>	+	water
e.g.	hydrochloric acid	+	zinc oxide	→	zinc chloride	+	water
	nitric acid	+	magnesium oxide	→	magnesium nitrate	+	water

## Displacement/competition

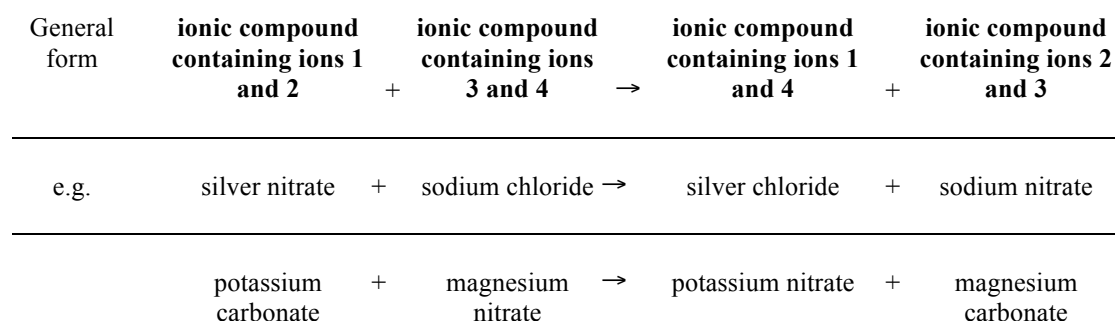
In these reactions a more reactive element ‘displaces’ a less reactive element from one of its compounds. The *metaphor* of ‘competition’ suggests that the more reactive element out-competes the less reactive element for the other element or radical (e.g. a nitrate ion).

General form	element 1	+	compound of element 2	→	compound of element 1	+	element 2
e.g.	magnesium	+	zinc nitrate	→	magnesium nitrate	+	zinc
	chlorine	+	sodium bromide	→	sodium chloride	+	bromine

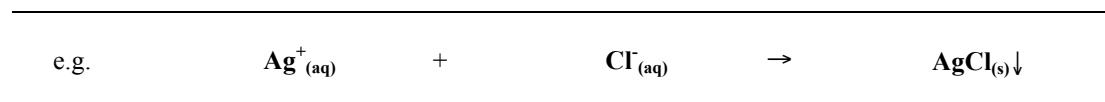
These reactions normally occur in solution, so the actual reaction concerns a redox process (see Chapter 7) involving two elements. In the examples above this would be (with the oxidation states, shown in brackets):



### Precipitation (double decomposition)



These are called precipitation reactions, because they take place in water ('aqueous' or 'aq') solution, and one compound is precipitated out of the solution (and can then be filtered, washed, re-crystallised etc). As with displacement reactions, some of the species present do not actually take any active part in the process (and so are called 'spectator' ions). So in the first example above sodium chloride and silver nitrate are both soluble, so what is actually being mixed is a solution containing hydrated sodium ions and hydrated chloride ions, with one containing hydrated silver ions and hydrated nitrate ions. The compound silver chloride has very low solubility (the bonding between the ions is not readily broken down to allow hydration of the ions), so is precipitated from the solution:



(It is worth noting that this equation makes it clear that ionic bonds form simply because of the attraction between ions: there is no need for a process of ‘electron transfer’ between the metal and non-metal. This is followed-up later in the Chapter.)

### **Chemical reactions and writing equations**

It is important to make it clear to students that being able to write an equation for a reaction, does not mean the reaction will happen (under observable conditions). So for example, as gold is a metal we can write an equation for its reaction with hydrochloric acid (hydrochloric acid + gold  $\rightarrow$  gold chloride plus hydrogen), but gold is an unreactive (‘noble’, rather than ‘base’) metal and should your school be in a position to keep a gold sample in the chemical stores, it would be safe from reacting with bench acid (normally 2 mol dm<sup>-3</sup>) at room temperature. Similarly, the equations for the displacement and precipitation reactions can readily be reversed, but the reactions only ‘go’ in one direction, determined by the energetics of the reaction at the conditions it is carried out (see Chapter 4).

### **Learning activities**

Some students will pick up the idea of general equations for type of reactions very readily, and will easily see the general forms as providing templates into which specific examples can be fitted. It will also seem obvious to many students how the general form of the equation (once learnt) allows us to see what reactions might be possible. So, for example, the reaction between sulfuric acid and potassium hydroxide is an example of the general type of reaction which produces a salt and water, and in this case the salt must potassium (from the particular alkali) sulfate (from the particular acid). However, research suggests that for many students, completing word equations is a mixture of guess-work, half-remembered ideas, and recollected patterns (some of which are inappropriate generalisations, or simply completely wrong).

This suggests that for most students, it is important to spend time on helping them appreciate the general patterns, and practice relating these general equations to a range of specific examples. As with many ideas in chemistry, it is also important to revisit whenever the opportunity arises. There are likely to be many potential opportunities for applying these ideas in a various topic contexts, provided a useful

basis for building-up student proficiency and confidence. The important thing is to remember that *once familiar*, these ideas seem very obvious, but it takes most students a good deal of engagement with using the ideas become they become familiar enough for this to be the case. Therefore it is important for teachers to be explicit in using these ideas – to model the logic clearly through speech when opportunities arise - until students are able to see the steps clearly themselves. It makes sense to introduce one general equation, and work through examples, before then gradually introducing other types of reactions so that students can effectively learn and discriminate between the different patterns. Scaffolding activities where students are given most of the information, but asked to complete examples and explain their reasoning can be differentiated for students with different levels of prior attainment, and made incrementally more difficult as students master the ideas. This can be the basis of group work, followed by a plenary where the teacher asks groups to explain the reasoning behind their answers, giving students opportunities to practice using the logic of relating general and specific equations.

### ***3.4 Modelling chemical change***

Chemical changes involve reorganisations of matter to produce different substances – and so involve both conservation and change. It was suggested above that it is important that students learn to appreciate what is conserved in a reaction, and what is modified.

Particle models, such as those discussed in the previous chapter, are important here, for chemists use structural models at the sub-microscopic level to develop explanations of chemical processes. These explanations often include consideration of:

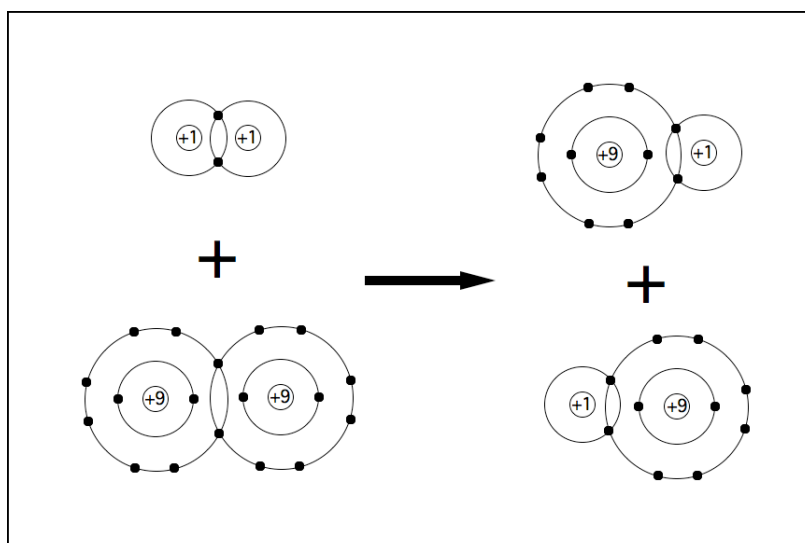
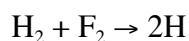
- the structure of reactants and products at the ‘molecular’ level;
- the relative stability / energy states of different species present before and after reactions;
- the reaction mechanism – a model of how the interactions between molecules and ions etc leads to change taking place.



These explanations often focus on the electrical structure of species – the way charge is distributed in a molecule, whether a species has a region of relatively low electron density where the positive atomic core can attract (and be attracted by) electrons on another species, and whether the existing charge configurations in species are effective at binding it together. A useful way of thinking about chemical change at the submicroscopic level, is to consider species such as ions, molecules and lattices as comprising atomic cores (a nucleus, usually surrounded by one or more ‘inner’ electron shells, see Chapter 1) and associated ‘valence’ electrons (which can be thought of as those from the outer or valence shells of the corresponding neutral atoms).

Consider, as an example, the reaction between hydrogen and fluorine:

hydrogen + fluorine → hydrogen fluoride



**Figure 3.3: Reactant and product species when hydrogen reactions with fluorine**

In this reaction the molecules of hydrogen and fluorine are broken up, and new molecules of hydrogen fluoride are formed. Matter is conserved at the level of the ‘quanticles’: the valence shell electrons that were around the hydrogen nuclei in the hydrogen molecules and the fluorine atomic cores in the fluorine molecules from the

original molecules are reconfigured around the atomic cores in the new molecules. As hydrogen is in the first period, so an atom only has one occupied electron shell, the atomic core in this case is just the nucleus. For elements in period 2 and above, the valence electrons surround an atomic core of nucleus plus inner shell(s) of electrons.

We can imagine that when we teach students about chemical reactions in the secondary school they might well wonder about why reactions occur. Yet this is not a question with a simple answer. A chemist looking to explain why this reaction occurs – at least under conditions where there is sufficient energy to break up some of the molecules – would need to find reasons why the molecules in the product species are more stable than those in the reactants. Another way of saying this is *why are the hydrogen nucleus and fluorine core binded together more strongly in HF molecules, than these species bind together in the molecules of the reactants*. A chemist might also want to suggest a reaction mechanism to explain how the reconfiguration proceeds – usually conceptualised in terms of the electrical interactions between the species present (and sometimes described in terms of the overlap of electron orbitals on different species). This level of thinking is clearly not accessible by most secondary students still familiarising themselves with our basic models of matter at the level of molecules and ions.

### **Avoiding the misconception that reactions occur to allow atoms to fill their shells**

Indeed, this is quite advanced material, and is usually only met by those who choose to continue with chemistry beyond the treatment in school science. It is certainly not usually considered suitable for inclusion in introductory chemistry classes. However, it is worth noting that by the time students do get taught about these ideas, they have already had years of observing and studying chemical reactions. Not surprisingly, they have usually developed their own ideas about what is going on amongst all these reacting molecules. Very commonly students who have successfully completed the chemistry in school science will explain this reaction in terms of atoms ‘needing’ to acquire full shells. They will argue (in relation to our example above) that a hydrogen atom only has one electron in its outer shell, and ‘needs’ two to be stable, and that a fluorine atoms has seven electrons in its outer shell, and ‘needs’ eight; and that by the hydrogen atom and fluorine atom sharing electrons they are both able to be ‘happy’.

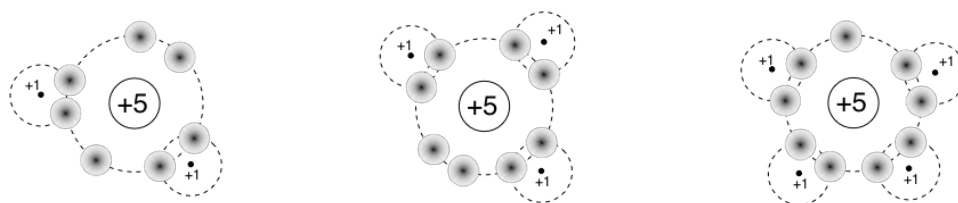
Students will often still give this type of explanation even after learning about the explanations that chemists develop in terms of energetics (see Chapter 4). Yet, of course, such an answer is not tenable, as it does nothing to explain why hydrogen molecules and fluorine molecules (which already provide the ‘full shell’ patterns students focus on) should interact and rearrange to allow the reaction to occur.

It seems important then, that although we usually consider the chemical ideas too complicated to teach to younger students, teachers should avoid encouraging the development of misconceptions that can later get in the way of progression in learning. The misconception that chemical reactions occur to allow atoms to fill their shells is very pervasive, and very tenacious, and teachers should avoid making comments that might encourage such ideas.

### **Proper use of the ‘octet rule’**

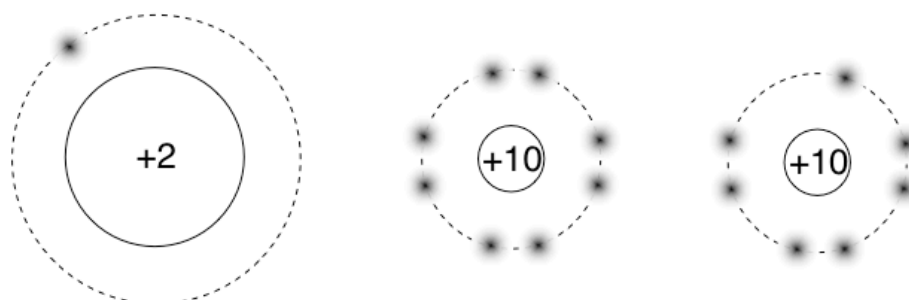
Even the idea that full shells are associated with chemical stability needs some careful presentation. The electronic structures of the species present in most stable substances tend to have particular patterns – and usually this means that there are eight valence electrons around each atomic core (apart from hydrogen where there are two valence electrons). This is a very useful rule of thumb, although there are many exceptions – stable substances where this pattern is not found

The ‘octet rule’ is very useful in highlighting which species are likely to be more stable when comparing ‘like with like’. So if a student is not sure whether nitrogen (electronic configuration, e.c. 2.5: that is two electrons in its first shell, 5 in its second shell) forms a hydride that has the formula  $\text{NH}_2$  or  $\text{NH}_3$  or  $\text{NH}_4$ , then the octet rule will correctly suggest the trihydride would be more stable. This is represented in figure 3.4, which shows that  $\text{NH}_3$  (but not  $\text{NH}_2$  or  $\text{NH}_4$ ) provides an octet of electrons around the central nitrogen atomic core (which comprising of a nucleus with charge +7 surrounded by a shell of two electrons is shown as having a +5 charge):



**Figure 3.4: Using the octet rule to identify the stable molecule**

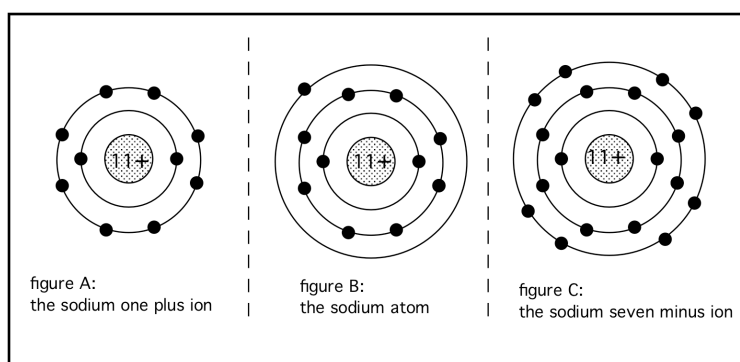
Similarly, if a student is not sure whether the common ion of magnesium is  $\text{Mg}^+$ ,  $\text{Mg}^{2+}$  or  $\text{Mg}^{3+}$ , then the octet rule tells us that Mg (electronic configuration, e.c.: 2.8.2) will form  $\text{Mg}^{2+}$  (e.c.: 2.8) rather than another ion (see figure 3.5). In these situations the rule is very useful. This is represented in Figure 3.5 (note how the  $\text{Mg}^+$  ion is shown as having a larger, but less charged, atomic core, as it has an electron in the third shell, unlike the  $\text{Mg}^{2+}$  and  $\text{Mg}^{3+}$  ions).



**Figure 3.5: Using the octet rule to identify the stable ion**

However the octet rule is only a heuristic, or rule-of-thumb, as there are plenty of exceptions: compounds such as carbon monoxide ( $\text{CO}$ ) and borane ( $\text{BH}_3$ ). Transition metals commonly form a range of ions (iron for example commonly exists in compounds as both  $\text{Fe}^{2+}$  and  $\text{Fe}^{3+}$ ). Most importantly, because most substances stable enough to be found in our normal surroundings, or indeed in chemistry laboratories, have electronic structures that already ‘obey’ the octet rule (as in the hydrogen and fluorine example above, Figure 3.3), it is of no help in explaining why they react.

Moreover, the stability of ions is always relative to the chemical environment. Sodium is a reactive metal that readily forms compounds containing the sodium ion,  $\text{Na}^+$ , so we think of  $\text{Na}^+$  as a stable species. This is a fair judgement, as the ion is stable *in normal chemical contexts*: in the metal lattice, in salts, in aqueous solution. However, what students often do not realise is that the (isolated) neutral atom is *more stable* than the (isolated) ion with its outer electron removed. Chemists actually do strip the electrons from atomised sodium (and other elements) to measure the ionisation energy: and energy is needed because the atom is more stable, despite not having a full shell. The removed electron will be attracted back if no other chemical species are around.



**Figure 3.6: Which of these sodium species do students consider stable?**

Ionisation of atomised metal may seem an extreme case, so perhaps it is not important if students think of the  $\text{Na}^+$  ion as more stable than the atom. However students have a strong tendency to see *any* species with an octet of electrons as stable, and research shows that by the end of secondary education, students will commonly rate a whole range of dubious ions as more stable than atoms because they have full shells or octets of electrons. So not only do students tend to think  $\text{Na}^+$  is a stable ion, they make the same judgement about the chemically quite ridiculous species  $\text{Na}^{7-}$  (see Figure 3.6). It is important, therefore, that teachers are careful to make sure that students do not over-generalise the octet rule from a very useful rule-of-thumb for identifying the most likely formulae for molecules and ions, and adopt it as an absolute principle to judge stability and explain why reactions occur.

### ***3.5 The chemical bond***

In considering chemical reactions, or indeed processes such as changes of state and dissolving, in terms of submicroscopic particle models, it is clear that the quantiles (ions, molecules etc) tend to clump together, and that the changes we observe at a macroscopic scale may be explained in terms of *reconfigurations* of those theoretical particles. We might think of these changes in terms of two types of competitive considerations: the effect of the attraction between quantiles working against the available energy associated with the quantiles (higher temperatures mean more energy for the quantiles to overcome their mutual attractions, cf. Chapter 2); and the effect of competitions between different possible arrangements (with chemical reactions occurring when there is *enough energy available* to disrupt one arrangement, and allow a reconfiguration into a more stable arrangement). This idea of ‘activation’ energy is discussed further in Chapter 4. A key concept here is that electrical binding holds together different quantiles to form the clumps.

#### **The nature of the chemical bond**

At the level at which we talk about bonding in secondary school chemistry, we can think of it as an electrical phenomenon. The basic components of matter are electrically charged (the nuclei containing positive protons, and negative electrons), and in chemistry it is often most useful to think of chemical change in terms of reconfigurations of positive atomic cores surrounded by negative electrons. Similar charges repel (atomic cores repel each other; electrons repel each other) but opposite charges attract (cores attract and are attracted by electrons), and the attractive forces pull the cores and electrons together into arrangements until similarly charged components are close enough for the repulsion to balance the attraction. This forms an equilibrium arrangement, which is stable because any small disturbance from this position will be resisted by the electrical forces. This applies to atoms, ions, molecules, lattices etc.

When students first meet the idea of the chemical bond, they have limited experience on which to construct a model in terms of electrical forces. Lower secondary students usually have limited appreciation of forces between charges, and they commonly

think of bonds as not just being physical (a force) but actually material. They may think in terms of glue, sticks, springs, and various other physical connectors – and some of these may well derive from the physical models we use in teaching (where plastic spheres are connected by such material means). It is difficult for students to think in terms of bonding that is not due to some material link, even if they can logically appreciate that at the level of electrons and ions and molecules, it is not possible for there to be smaller material links holding them together. Handling this as a teacher requires some subtlety. When teaching abstract ideas, it is important to find connections with students' existing experience and ideas, however the notion that the chemical bonds are material links is clearly a serious misconception. As teachers we need to welcome students making comparisons with what is familiar – the bond *acts like* glue, or *like* an elastic band joining two things together – whilst trying to ensure that this analogy does not become seen as an identity (e.g. not that the bond *is* a spring between two atoms).

Some students may think of the bonding in the atom as being like that due to magnets. This is not a perfect comparison (as magnets have two poles, unlike charged particles), but students will be familiar with magnetism acting as a physical force that can occur without material connection, and in this sense this may be a more appropriate analogy. However it remains important to emphasise that the bonding is electrical, and acts *like* magnetism. (At higher levels of study those students who continue with chemistry will find that actually there is a magnetic component to atomic structure and bonding, but at the introductory level this is an unhelpful complication. Indeed, ultimately, magnetism is an electrical phenomenon, but again that is an unhelpful complication here.)

### **Quantum theory**

The electrical model is not the whole story, as it does not explain why the stable arrangements produced by the electrical forces so often result in species with particular electronic configurations (two electrons in the first shell, and often eight in the outermost shell of a species). This is not normally tackled until more advanced levels of study, as it is usually considered to be a difficult topic. In principle, however, this relates to how at the tiniest scale, everything is quantised: such quantities as

energy and angular momentum are found to exist in minimally-sized packets. This is very similar to the way matter and charge are quantised, giving us basic units such as electrons that cannot be divided into anything smaller. That also seems a very counter-intuitive idea, but one that students usually come to accept given plenty of opportunity to explore and use our particle models of matter.

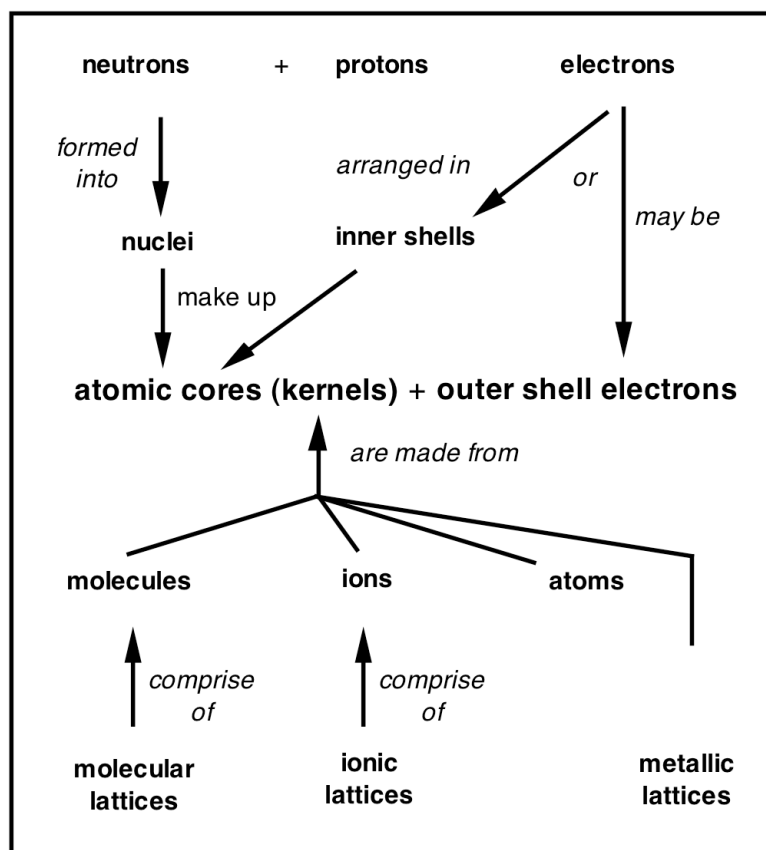
A strong clue that the electrical model cannot be the whole story is how in atomic and molecular structures, electrons are often considered to be found in pairs – pairs of electrons form covalent bonds between atomic cores, and non-bonding pairs of electrons on some atoms form hydrogen bonds with hydrogen atoms bonded to other atoms (see below). If electrical forces were the only important factor, then electrons would not be expected to act as paired-up. Yet students rarely seem to raise this as an issue, something that indicates how difficult it is to persuade them to think primarily in terms of electrical interactions. A teacher who finds students raising this objection can probably consider they are doing better than most in helping students think about matter at the submicroscopic scale in the way chemists do.

(The electrons do repel, but the existence of quantum mechanical spin can in effect reduce this effect for pairs of electrons. One model of this is to think of electrons as tiny magnets which can be arranged anti-parallel - N-S and S-N - so that they have a magnetic attraction to counter the electrical repulsion.)

### ***3.6 Modelling varieties of chemical bond***

At a basic level, the different types of bonds discussed in chemistry can all be understood as due to the electrical interactions between different species, attracting them together until the repulsions between similar charges balance the attractions between opposite charges (see Figure 3.7). There is often a net attraction between neutral species such as molecules, for example, until they get very close when this is balanced by repulsions. This is an essential prerequisite for matter as we know it to exist – otherwise the universe would probably either have no condensed matter, or would be one large neutron star!





**Figure 3.7: A way of thinking about the structure of matter at submicroscopic levels**

There is some ambiguity among chemists and in different textbooks over whether the term ‘chemical bond’ should be applied to all or just some of these interactions:

usually agreed to be chemical bonds	not always seen as chemical bonds
covalent (including dative)	intermolecular interactions (such as van der waals’ forces)
hydrogen	solvent-solute interactions
ionic	
polar	
metallic	

Such a distinction is usually based on considering those interactions labelled as bonds as being stronger than the others. However, that is not a clear distinction, as we see when something like NaCl (with strong chemical bonding) readily dissolves in water

– where the strength of the interactions between the ions and the polar water molecules are a major factor in its solubility.

Even the term ‘intermolecular’ bonding needs to be applied carefully. Hydrogen bonding (usually considered to be a chemical bond, although sometimes that status has been questioned) can be intramolecular (being very important in determining the shape of proteins and nucleic acids, for example) or intermolecular.

Dative bonding *could* be considered intermolecular as well, so if gaseous ammonia (NH<sub>3</sub>) and hydrogen chloride (HCl) are mixed, they form a solid (NH<sub>4</sub>Cl). This may be considered to be an ‘adduct’, the name given when two already stable molecules are able to join into a new larger molecule (i.e. NH<sub>4</sub>Cl might also be written as NH<sub>3</sub>.HCl). If the product here, ammonium chloride, is gently warmed then it thermally decomposes back to hydrogen chloride and ammonia.



The formation of the ammonium chloride is easily demonstrated in the lab: a long glass tube with cotton wool at either end can be used (see Chapter 2). As little energy input is needed to bring about the decomposition, there is a tendency to see the bond formed as intermolecular, and consider the adduct as not being a fully stable molecule. Clearly such complexities are unhelpful when introducing the topics of chemical bonding to students. Moreover, individual students have been found to use terms such as bond, bonding, attraction etc., in their own idiosyncratic nuanced ways.

It is more sensible to consider all of the electrically based interactions that hold together molecules (internally), lattices, and complexes such as adduct and solvated species as forms of chemical bond, and make it clear that some bonds are much stronger than others, and so much more energy is needed to disrupt the stronger bonds.

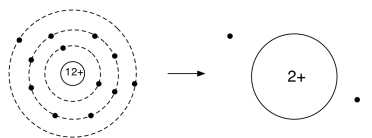
## **A recommended teaching sequence**

Students are prone to see all materials as made up of molecules, and in particular to see ionic compounds as containing discrete molecule-like entities. This tendency is encouraged when introducing covalent bonding first. Here it is recommended that you take metallic bonding as a starting point for thinking about bonding, moving on to first ionic, then covalent bonding. It is important to be explicit that what you are teaching are *models* that chemists use to explain properties of different substances. These ‘first-order’ models offer a good deal of explanatory power, but do not explain everything, and do not fit all cases.

## **Metallic lattices**

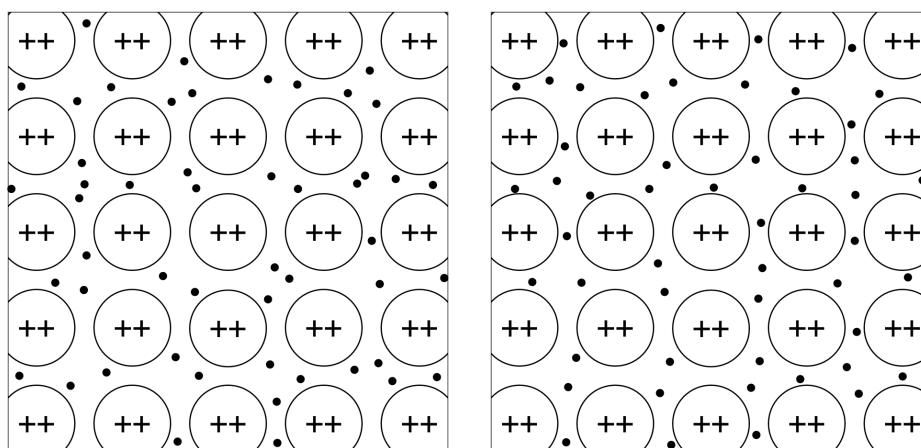
Metals are elements, and so having relatively simple structure provide a good place to start modelling bonding. At the simplest level, the structure of metals can be modelled as a regular arrangement of atomic cores, that is metallic cations, which despite all having a positive charge are held in their lattice positions because of a large number of electrons able to move around and between them, acting as a kind of ‘electrical glue’. This is of course a metaphor, and should be used carefully.

A very common metaphor used to describe this arrangement is that the electrons form a ‘sea’ in which the cations are immersed. However, students using the ‘sea of electrons’ idea sometimes understand it to mean that there is a *vast excess* of electrons, whereas the stability of the metallic lattice depends upon its overall neutrality. In teaching this model it is important to stress that the number of electrons per cation is the same as the magnitude of the core charge, as the overall neutrality of the lattice is an important factor in its stability. So in the example of magnesium (see Figure 3.8), the electronic configuration of the atom is 2.8.2, which can be modelled as a positively charged atomic core, plus two valence electrons.



**Figure 3.8: magnesium atom as an atomic core ( $\text{Mg}^{2+}$  cation) plus two valence electrons**

The metallic lattice consists of a vast array of these cores,  $\text{Mg}^{2+}$  cations, with the associated electrons. We can think of the lattice forming when a vast number of magnesium atoms come close enough together for their outer shells to overlap and merge so the electrons in them can move throughout the array. This is shown in Figure 3.9, where a sectional slice is shown – with the electron positions shifting from one moment to the next. It is useful to also use three-dimensional models to emphasis that the structure is not just ordered in two dimensions.

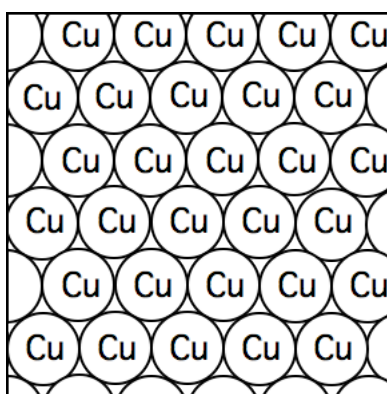


**Figure 3.9: Metallic bonding holds together the lattice in metals such as magnesium – the bond comprises of the mutual attraction between the metallic atomic cores, and the delocalised electrons able to move around the structure**

A simple way to reinforce the electrical neutrality of the metallic lattice would be to present students with a series of images (similar to those in Figure 3.9) showing core charges of +1, +2, and +3, some showing a balance of charge, and some with substantially too few, or too many electrons. Students could be asked to work in

groups to identify which figures represent plausible structures, and to justify their decisions. This activity might be especially suitable for less highly achieving students.

The simple arrangement in Figure 3.9 reflects a ‘cubic’ arrangement of cations. This is just a model. Although some metallic structures are cubic, most are based on each ion in a slice of the lattice being surrounded by six others in a hexagonal arrangement (see figure 3.10), as this actually gives better packing (‘close’ packing) when the next layer is offset.



**Figure 3.10: A representation of close-packing in copper**

Modelling this with spheres (such as marbles, or expanded polystyrene balls) shows that there actually two regular ways of building up such layers, depending upon whether the third layer sit directly above the first (‘ABAB’), or is offset from both the first two layers (‘ABCABC’). This can provide the basis of simple practical modelling activity for students to build models of the two different arrangements using suitable spheres (such as expanded polystyrene balls one size). Although secondary students are not usually expected to know about these specific structures, the modelling activity provides a suitable group practical that can reinforce teaching about the regular nature of the metallic lattice. With some students, close instructions for building the model will be appropriate, whereas for others they could simply be issued the challenge of producing two non-identical structures with all the spheres close packed.

The apparent ambiguity between figures such as figure 3.9, which shows space between the atomic cores for the electrons, and physical models showing the close packing provides an important teaching point about the difficulty of modelling

quanticles – atoms and ions do not actually have clearly defined surfaces of boundaries, but rather become more tenuous further away from the nucleus.

A more advanced way of thinking, usually only met at college ('sixth-form') level explains metallic structure in terms of a more complex orbital model, where the metallic bonding is formed by the overlap of atomic orbitals, and the outcome is an enormous number of 'molecular' orbitals that will have a complex pattern of geometries. However, they will also form a virtual continuum of energy levels (the 'conduction band'), so although each particular molecular orbital may put restrictions upon occupying electrons, the available thermal energy is sufficient for electrons to readily move between orbitals in the band.

Although this is a more complex picture, and will only met by advanced learners, it is important to ensure introductory teaching will not act as an impediment to later progression for those students who do continue with the subject. So, in teaching in terms of overlap *of shells* we should be careful to stress that this is a model, and somewhat simplified, so that students who may study chemistry at higher levels do not become too committed to that particular picture. Given this proviso, the overlapping shells model can act as a much simpler 'version' of the molecular orbital/conduction band model to explain the delocalisation of the valance electrons – the 'conduction' electrons so important to the properties of metals.

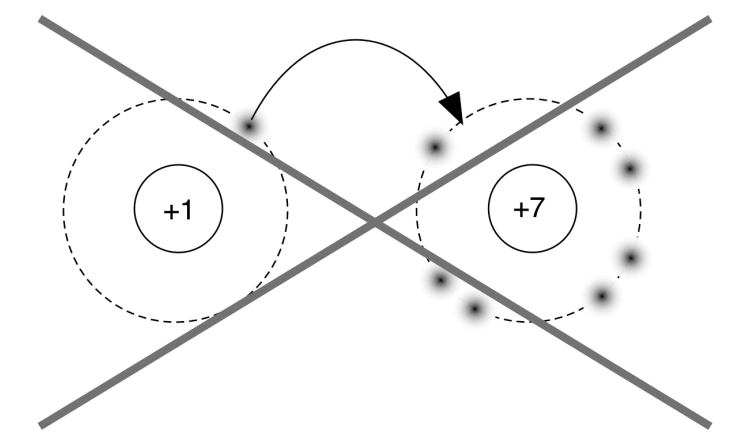
The key points when introducing metallic bonding is that the cations form a regular pattern, and are bound by the attraction between the positive cations and the negative electrons. As always, our diagrams need to be presented to students as representations designed to emphasise certain points, and not as realistic images of how metals actually are.

### **Ionic lattices**

Somewhat more complicated than the metallic case, is that of ionic compounds. Research shows us that students very commonly misunderstand ionic bonding. Ionic bonding is often taught through a convention of considering atoms of an electropositive metal and an electronegative non-metal - often Na and Cl are used as

the example - and then considering how they might interact to form ions,  $\text{Na}^+$  and  $\text{Cl}^-$ , which would then bond together (see Figure 3.11).

This does not reflect a likely chemical process, but is just a kind of ‘thought experiment’ in how ions might be formed from atoms, in the unlikely event that such unstable species should be around to happen to interact. However, students often think such a scheme is actually how ionic bonds are formed, and - worse - often think that the electron transfer depicted in such schemes ‘is’ the bond!

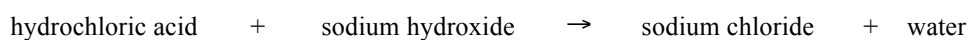


**Figure 3.11: A common student misconception of the ionic bond**

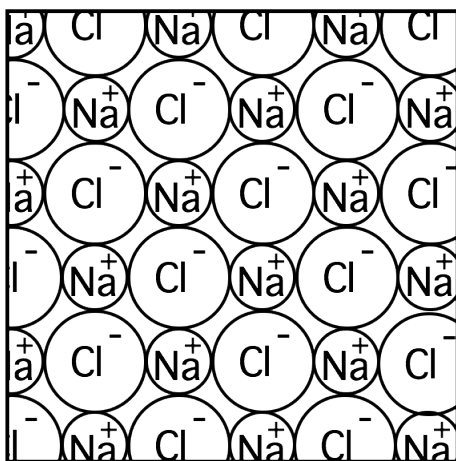
However, we *do not have to consider how ions come about to explain ionic bonding*: indeed presenting such schemes only encourages students to think of all chemical processes as starting with atoms, rather than more feasible reactants. In the natural world there are many materials that already contain ions like  $\text{Na}^+$  and  $\text{Cl}^-$ , but we never find *atomic* sodium or chlorine under natural conditions. (Indeed, there are very few materials that contain discrete atoms: samples of the noble gases being the obvious exceptions.) In the case of metallic bonding it was useful to start thinking about atoms, but thinking in terms of atoms in explaining ionic bonding is an unhelpful mind-set.

It is recommended that in introducing ionic bonding to students, it is more useful to think in terms of a more feasible chemical context, such as in terms of reactions they

are expected to be familiar with. One example might be the neutralisation of an acid and an alkali, such as



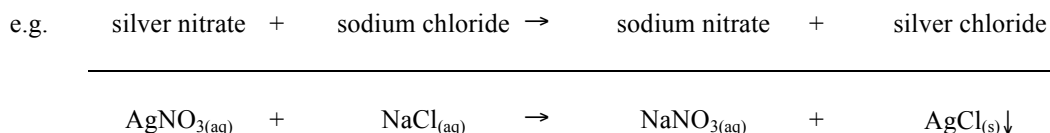
In this case the reactants are solutions containing ions, and the actual chemical reaction is between hydrogen ions and hydroxide ions to form more water. This leaves sodium and chloride ions in solution:  $\text{Na}^+_{(\text{aq})}$ ,  $\text{Cl}^-_{(\text{aq})}$ . Solid sodium chloride does not form because the ions are too strongly hydrated (they are each bonded to a sheath of solvent molecules, which forms an ad hoc complex in the solution). However, if the solvent, the water, is allowed to evaporate (see Chapter 1), then this leaves the ions, which organise into a regular array of cations and anions because of the mutual attraction between oppositely charged ions (see figure 3.12).



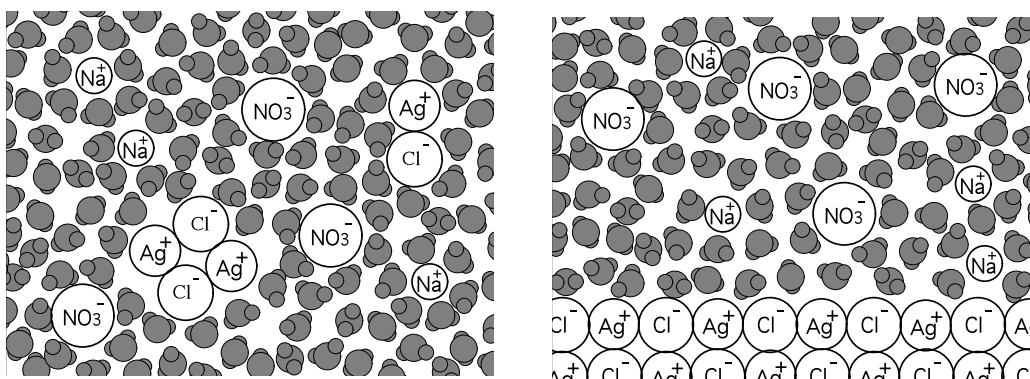
**Figure 3.12: Representation of the ionic sodium chloride lattice in two dimensions**

An alternative context for introducing ionic bonding would be a double decomposition reaction, because the students can see the formation of the ionic product immediately. So if silver nitrate solution is added to sodium chloride solution, the formation of the precipitate, silver chloride, is immediate.





It is important to stress to students that although we call the reactant solutions ‘silver nitrate’ and ‘sodium chloride’ solutions (as they are solutions of these compounds), the solutes do not exist as bonded compounds in the solution, but rather as ions which are mixed into the solvent. However, when the two solutions are mixed together so that it contains silver, sodium, chloride *and* nitrate ions, the attraction between silver and chloride ions is strong enough that when they collide in the mixture they bind together, eventually forming large clumps that settle from the solution to give a solid silver chloride ‘precipitate’ (see figure 3.13).



**Figure 3.13: When solutions of silver nitrate and sodium chloride are mixed, the silver and chloride ions bond together forming a solid with an ionic lattice.**

In this process an ionic lattice is formed because of the attraction between the oppositely charged ions. In essence, that *is* the bond. There is no need to explain how the ions came to be – they exist in materials available to chemists, in substances we call salts (NaCl, NaBr, KCl, K<sub>2</sub>SO<sub>4</sub>, Ca(NO<sub>3</sub>)<sub>2</sub> etc). Most of the material on earth comprises of the output of nuclear processes that took place in stars, where the temperature is much too high for individual atoms to exist (and so matter is in the form of plasma, a kind of gas comprising of separate nuclei and electrons). After this material was exploded into space, at the end of the star’s ‘life’ cycle, it cooled and

formed into more stable combinations of nuclei, and electrons: sometimes atoms, but sometimes ions or molecules. Very little of the matter that formed the earth is of the form of discrete atoms. The formation of the ionic lattice due to ionic bonding does not involve any mysterious ‘electron transfer’, and if ionic bonding is taught this way there is no need for students to think in those terms.

Of course the formation of sodium chloride by binary synthesis *does* involve the formation of chloride ions (arguably the sodium ions are already present in the metallic lattice), but from molecules, not individual atoms. Students may see this reaction demonstrated, but they are unlikely to carry it out (on safety grounds); and neither metallic sodium nor chlorine gas are common laboratory reagents. Whilst the binary synthesis route certainly offers an exciting demonstration of a vigorous reaction, forming sodium chloride by neutralisation followed by evaporation is much more practicable way of producing sodium chloride.

This is important because research tells us that students often think that in the NaCl lattice, for example, there are NaCl molecules, or at least discrete ion-pairs, which are bound because they have a history of having transferred electrons: so despite the symmetry of Figure 3.12, students often interpret such figures as a collection of NaCl molecules which have ionic *bonds* within them, but are then just attracted to each other *just by forces*. This is unfortunate as such misconceptions (a) cannot help students understand why NaCl is hard, and has a high melting temperature, and (b) often lead them to expect NaCl molecules to be the solvated species when a solution is prepared. These misconceptions all seem to derive from teaching the ionic bond through fictitious electron transfer events between isolated atoms, which actually are quite irrelevant to the chemistry! This idea appeals to students so much that they will sometimes explain precipitation reactions, such as our AgCl example in terms of

- the silver ion getting its electron back from the nitrate, and
- the chloride ion giving back the sodium atom its electron, so that silver and chlorine are back to being atoms;

- allowing the silver atom to then give an electron to chlorine atom, to reform the silver and chloride ions (which were already present, of course) with an ionic bond between them.

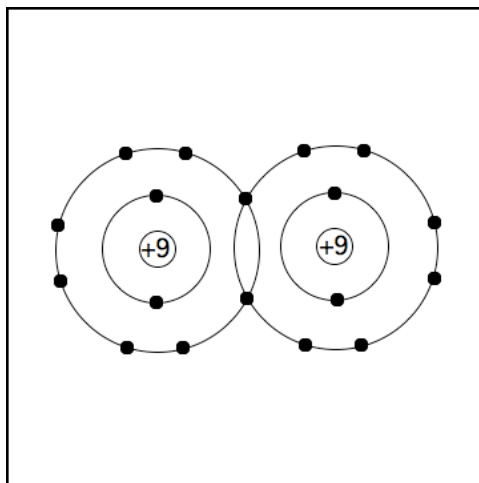
This scheme is actually much more complicated than the scientific model of silver ions sticking to chloride ions because of their opposite charges, and illustrates just how tenacious some misconceptions can be, once they have a hold of a student's imagination.

### **Covalent bonding**

Covalent bonding tends to occur between non-metallic elements, and is often described using the metaphor of 'sharing' electrons. A covalent bond is understood as occurring when the valence shells of two atoms overlap so that one (or more) pairs of electrons fall within the valence shells of both atoms. At more advanced levels this is described in terms of the interaction of atomic orbitals on different atoms forming molecular orbitals (and the pair of electrons occupying the lower energy 'bonding' molecular orbital), but at an introductory level, atoms are often represented as having overlapping outer electron shells.

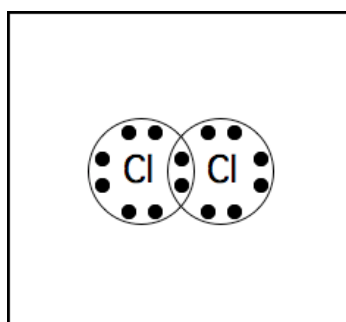
We can represent covalent bonds in a variety of ways, which can be confusing for students. Experienced chemists and teachers see past these differences in representational formalism, but the reasons for different ways of drawing the same thing may seem arbitrary to learners. Indeed, when looking at a range of student textbooks there is often no obvious reason for the preferred forms of diagrams used. These images represent 'quanticles' – entities that are fuzzy and often better thought of as clouds of charge than as like tiny billiard balls with definite surfaces. This may not be what our common forms of representation suggest (for example, drawing electron shells may give the impression that they are solid structural elements of atoms), so it is up to the teacher to emphasise the limitations of images which are often a compromise between what we can easily draw, and particular points we wish to represent in particular images.

Some examples of how covalently bound molecules may be represented are illustrated here. Figure 3.14 shows a molecule of fluorine where the covalent bond is represented as the pair of electrons where the outer shells overlap.



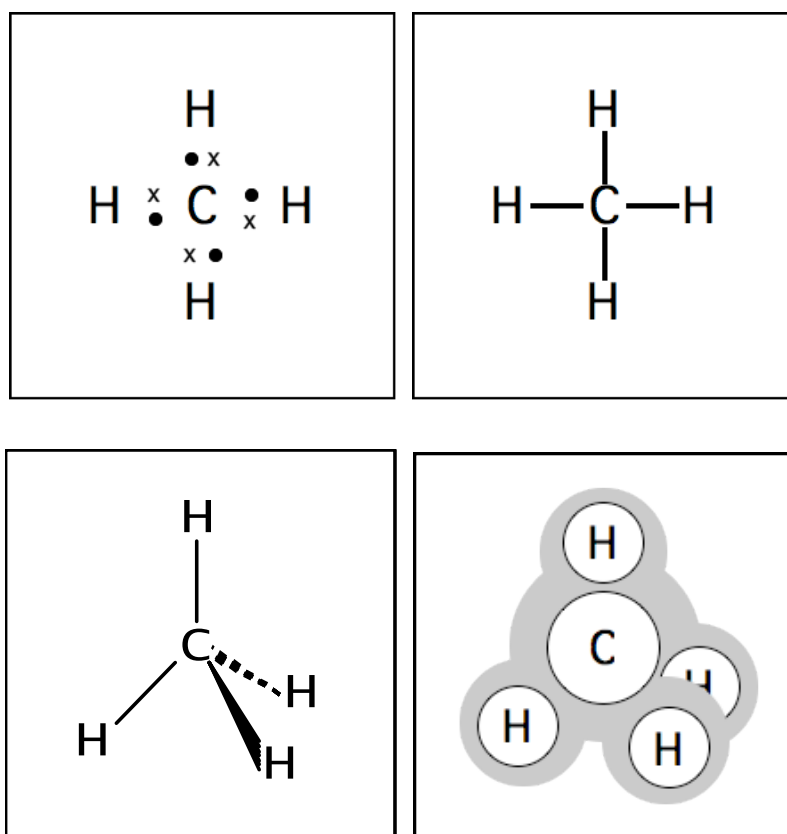
**Figure 3.14: One possible representation of a molecule of fluorine (F<sub>2</sub>)**

In figure 3.15, the chlorine molecule is represented in an alternative way, with only the valence shell electrons shown, and the bonding pair of electrons shown inside the overlapping atoms. It is not sensible to ask which of these pictures is a more accurate representative of the molecule (which is too small to be visible, three dimensional, with electrons in motion) itself, but rather it is important to explain to students which features are being foregrounded in different forms of representation. In terms of the covalent bond itself, the essential feature is that the pair of electrons is electrically attracted to, and by, the positive nuclei, and so act as a bond.



**Figure 3.15: One possible representation of the chlorine (Cl<sub>2</sub>) molecule**

Figure 3.16 shows some alternative ways of representing the methane ( $\text{CH}_4$ ) molecule. The first of these uses a ‘dot and cross’ style of showing valence shell electrons. This can be helpful for students in *keeping account* of electrons by showing the electrons from different atoms in a different style. However, in the molecule, there is no difference between the electrons, and the interactions between the electrons and the positive nuclei are completely independent of where the electrons derive from. This should be stressed, as some students assume that each electron in the bond is more strongly attracted to its ‘own’ atom, and that on bond fission, the electrons will always go back to the atoms they came from (which will interfere with later learning about heterolytic bond fission). Students should be nudged from thinking in terms of the ‘ownership’ and ‘history’ of electrons, to instead thinking simply in terms of the electrical forces acting between the different charges present.



**Figure 3.16: several representations of the methane ( $\text{CH}_4$ ) molecule**

The second representation in Figure 3.16 simply shows the bonds as lines connecting the parts of the molecule together. This is an easy representation to draw (which can

be important with more complicated molecules), but is probably best only used once students are used to the idea that the bond is the electrical interaction due to the electrons found between the different atomic nuclei. The third representation shows how the molecular structure is not flat like the surfaces where we draw, but takes up a three dimensional arrangement (tetrahedral in this case, due to the mutual repulsion of the electron pairs in the bonds). The final image replaces the representation of discrete bonds, with a representation of the ‘clouds’ of electron density around the atomic cores – i.e. how we can imagine a time-averaged image of where the electrons are to be found. (From a quantum mechanical perspective, the structure of molecules is better described in terms of the probability of finding electrons in particular positions, but it is useful to think of this as the electrons moving about and so smearing out their charge density.) A more sophisticated version could show the variations in electron density – more like a contour map.

As I have suggested throughout the chapter, it is not helpful to talk about *which* of the various possible representations is ‘best’. Students should be made aware molecules and other ‘quanticles’ are not easily drawn, and that scientists will model them through representations that stress particular relevant features. There is a key issue here for teachers, as it usually makes good sense to adopt particular conventions in teaching and then to consistently use them to limit learning demand for students. Yet figures students see in books and on-line will be diverse, and reliance on one form of representation can lead to students treating that form as a realistic image of how molecules actually are. A sensible compromise would seem to be in order: that is initially using a preferred form to represent molecules but later (once students are used to seeing and drawing images of molecules) introducing variation where context makes others forms useful to make particular teaching points (e.g. about molecular shape or the presence of double bonds).

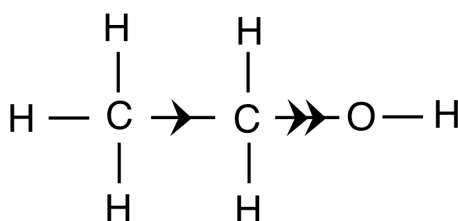
### **Bond polarity**

At an introductory level, students tend to be given the impression that bonding in compounds is covalent or ionic, as if this is a dichotomy:

Covalent bonding	Ionic bonding
In non-metallic elements and compounds of non-metals	In compounds of metals with non-metals

Students often come to see these different forms of bonds as fundamentally very different, making it difficult for them to later appreciate how few bonds are ‘pure’ covalent and indeed no bonds are ‘pure’ ionic.

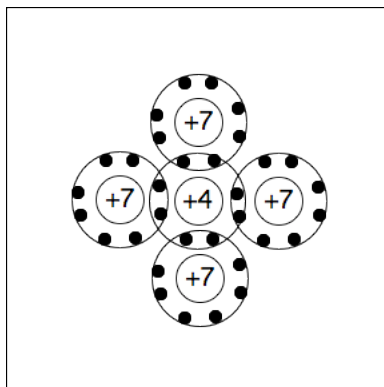
The ionic bond as represented in introductory chemistry texts is an ideal, and most compounds thought of as ionic are actually somewhat from having fully ionic bonding. (At advanced levels, students will learn how tables can be used to estimate the percentage of ionic and covalent character, depending upon the electronegativity difference between the elements.) Fully covalent bonds only usually exist between atoms of the same elements, and strictly then only where those atoms are not themselves bonded to very different atoms – so the C-C bond in ethanol is not pure covalent, as one of the carbon atoms is bonded to an electronegative oxygen atom, which will influence the carbon-carbon bond through an ‘inductive’ effect, distorting the geometry of the electron density in the carbon-carbon bond (Figure 3.17).



**Figure 3.17: The C-C bond in ethanol is not a ‘pure’ covalent bond, because the oxygen atom (core charge +6) attracts electron density and distorts the bond indirectly.**

These issues are ignored in introductory treatments, but some options for representing molecules offer a better starting point for later progression in students’ thinking. So Figure 3.18 shows a representation of a tetrafluoromethane (CF<sub>4</sub>) molecule showing the core charge and valence electrons. If students are taught about bonding as an electrical interaction, it seems clear that the electron pairs in the bond will – all other

things being equal - be pulled closer to the +7 charge (fluorine atomic core) than the +4 charge (carbon atomic core). So the bond here will be polar rather than purely covalent.

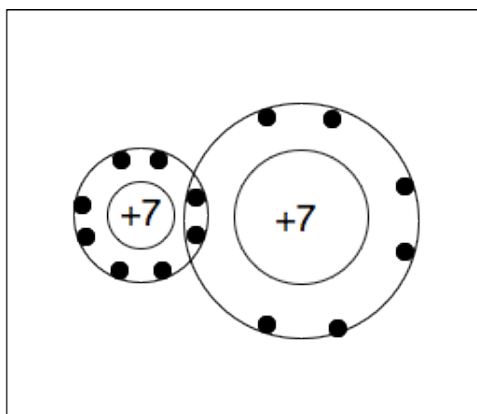


**Figure 3.18: a simple representation of the tetrafluoromethane ( $\text{CF}_4$ ) molecule**

Figure 3.19 shows a representation of an interhalogen compound, ClF, using the same format. Here both halogen atoms have the same core charge, +7, but because of the difference in size of the two cores (for fluorine, nucleus {+9} and one shell of 2 electrons; for chlorine, nucleus {+17} and two inner shells {2.8}), the equilibrium position for the bonding pair of electrons is nearer the fluorine nucleus. So again, this is a polar bond.

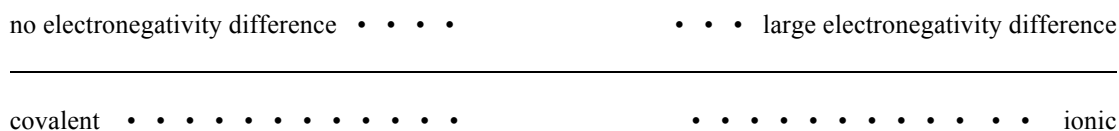
At advanced levels there are various ways of showing bond polarity (electron position, the use of  $\delta+$  and  $\delta-$  symbols to indicate 'partial' charges on atomic centres etc.) Whilst such detail is not needed in introductory treatments, it is important to teach bonding as primarily an electrical interaction, to help students later appreciate how ionic and covalent bonds can be understood as extremes on a continuum, and not a simple dichotomy where all bonds in compounds easily fit one or other category.





**Figure 3.19: a representation of the Cl-F molecule**

The extent to which a bond will be polar then depends on the difference in electronegativity of the elements involved:

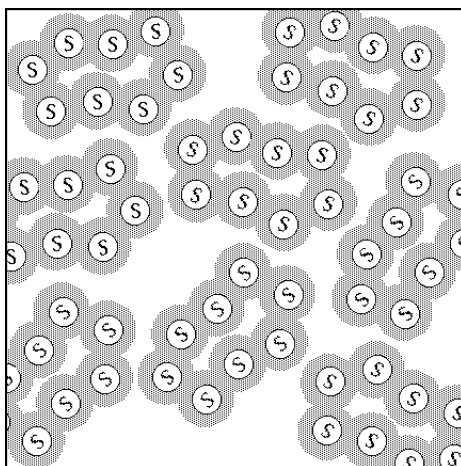


**Other forms of bonding**

In secondary chemistry it is common to limit explicit discussion of bonding to the metallic, ionic and covalent cases. However, students will have come across the idea that substances in the solid state are held together by some form of bonding when learning about the basic particle model of matter (see Chapter 2) and so are likely to assume that this bonding will be ionic in ionic compounds, metallic in metals *and covalent in material with covalent bonds*. The latter assumption would be correct in carbon, silicon and other substances with giant covalent lattices. However, there are many materials with discrete covalent molecules, which exist in the solid state at room temperature because of the weaker interactions between molecules. For example, wax and polyethylene (polythene) contain molecules that have covalent intra-molecular bonding, but are attracted to each other by a weaker form of (inter-molecular) bonding.

So Figure 3.20 shows the molecules in sulfur ( $S_8$ ), which is in the solid state at room temperatures, and can be melted by heating in a test tube over a Bunsen flame (care if

this is demonstrated – some of the sulfur may burn, leading to noxious vapour being released).



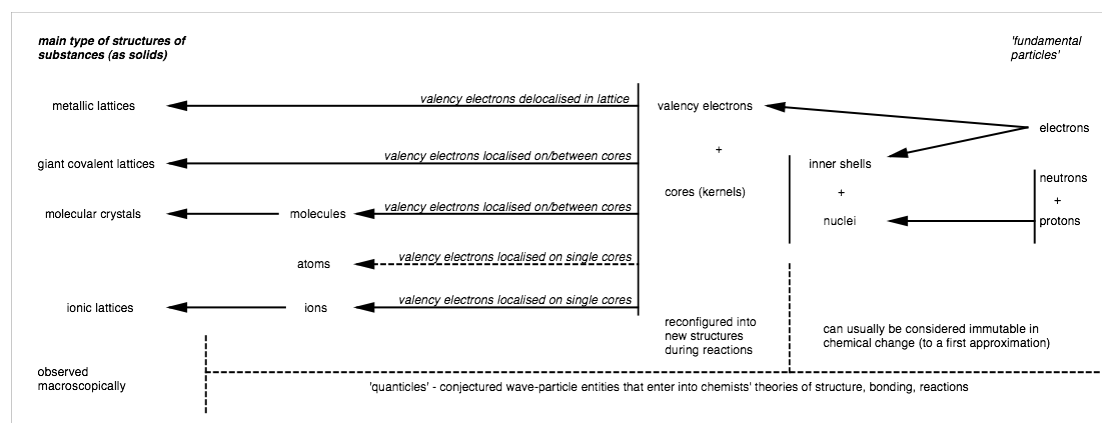
**Figure 3.20: Molecules in sulfur**

Sulfur comprises of molecules in the form of rings of eight atomic centres, with each atomic core bound to two neighbours by a covalent bond. However, sulphur exists in the solid state at room temperature because the molecules are attracted to each other (and actually fit together to give a crystalline structure). This is despite there being no electron ‘sharing’ *between* molecules, or any ions present. This is normally explained in terms of the electron movements in adjacent molecules becoming synchronised to give ‘fluctuating transient dipoles’ that allow areas of higher electron density (overall negative charge) on one molecule to attract and be attracted by areas of lower electron density (overall positive charge) on an adjacent molecule.

This is difficult to visualise, and is not usually discussed with students in introductory chemistry. However it is important students appreciate that neutral molecules will attract together because of the charges present, even if no details are offered. This will help to avoid the common misconceptions that so-called ‘molecular solids’ have covalent bonds throughout, and the corollary that, as many of these substances melt readily when in the solid state, covalent bonds are often quite weak.

Students need to appreciate that bonding effects are generally explained in terms of interactions between charges (and it may be worth pointing out that the forces attracting molecules of sulfur together can not be explained in terms of forming octets or full shells!) Figure 3.21 develops this principle (summarised earlier in Figure 3.7)

by showing how different chemical structures are built up of different configurations of atomic cores and valence electrons. Although it is certainly *not* sensible to present such a scheme to be learnt by students first meeting bonding ideas, it can usefully inform teaching. Teaching consistent with this way of thinking is more likely to help students appreciate the principles common to different forms of bonding, and will better support progression for those who go on to more advanced study.



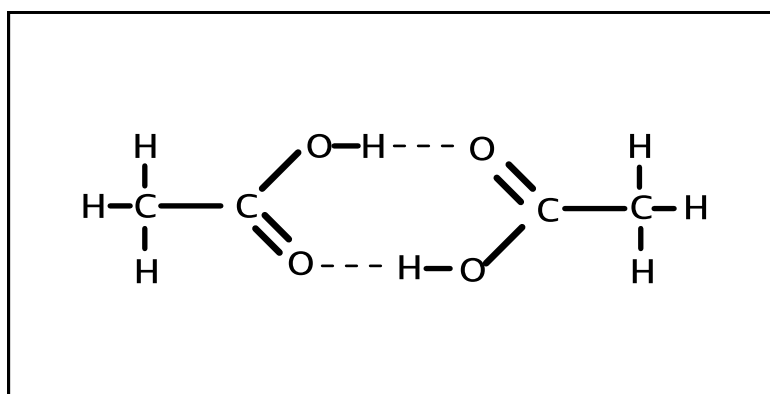
**Figure 3.21: Macroscopic structures reflect different ways of binding atomic cores together with valence electrons**

These types of interactions, often called van der Waals' forces, are not the only important types of bonding beyond metallic, ionic and covalent bonds. Solvent-solute interactions may be due to transient dipoles, but can also often involve permanent (rather than just transient) polarity on molecules. This is why water is a good solvent for ionic materials.

An important example of an interaction due to bond polarity, is the hydrogen bond, which is so important for the properties of water, protein and nucleic acids.

Unfortunately, students sometimes come across hydrogen bonding discussed in biology lessons (where the nature of the bond may not be explained) before it has been introduced in chemistry, leaving them to infer what is being referred to. (A common guess in this situation seems to be that it is just a covalent bond to a hydrogen atom.) Hydrogen bonds form between electronegative atoms (usually O, F, N, sometimes S, Cl; all of which have one or more pairs of non-bonding electrons, so called 'lone pairs') and hydrogen atoms that have polar bonds to other electronegative

atoms. Figure 3.22 shows hydrogen bonding between two molecules (a dimer) of ethanoic acid (as found in ‘glacial’, i.e. solid, ethanoic acid).



**Figure 3.22: hydrogen bonding in ethanoic acid**

Hydrogen bonds can be quite strong for intermolecular bonds (as a hydrogen atom bonded to an electronegative atom is a positively charged proton which has very limited electron density covering its ‘rear’ and so is readily attracted by/to ‘lone pairs’ of electrons on other atoms.) Researching into the nature of the hydrogen bond could be a useful extension activity to differentiate for more able (‘gifted’) students in a mixed-ability class. For example, hydrogen bonds have a specific geometry (the hydrogen bond being approximately opposite the polar bond), whereas most intermolecular bonds are non-directional: something that could be followed upon by a gifted student by exploring orbital models of bonding. A small group of more able students could be set the task of building a model of ice structure (or even part of a nucleic acid molecule) showing the importance of the intermolecular bonds in the crystal structure.

### ***3.7 Relating bond type to changes in chemistry***

Sometimes the weaker forms of interaction, such as van der Waals’ forces, solvent-solute interactions, and even hydrogen bonding, are considered to not really count as chemical bonds. However, as the strength of most types of interactions vary considerably, it makes more sense to think of there being a range of different types of bond, some of which are *usually* stronger, and some of which *tend* to be weaker. So

sulfur, held together by van der Waals' forces, has a higher melting temperature than mercury, although the latter has metallic bonding.

## Melting and boiling

A simple change of state, such as melting, has different consequences for substances depending upon the type(s) of bonding present.

Metals	<p>Metallic bonds are not strongly disrupted by melting (e.g. the metallic bond is not dependent upon a particular geometry, so the metal is still held together by metallic bonding in the liquid state, and – for example - continues to conduct electricity), and some metals have quite modest melting temperatures (e.g. Na c.98°C), but boiling requires overcoming the metallic bonding completely (so sodium has a boiling temperature of 890°C). Group 2 metals have higher melting temperatures than group 1 metals (as would be expected from having greater core charges and more delocalised electrons).</p> <p>Transition metals tend to have higher melting temperatures than main group metals, and their bonding is said to include some 'covalent character'. (Here our model that bonding only involves the outermost shell of electrons is found to have limitations.)</p> <p>Boiling a metal is basically a form of atomisation – although metal vapours may also include small clumps of atoms.</p>
Ionic solids	<p>Ionic solids tend to have high melting temperatures, as the ionic bonding depends upon the ions being arranged in the lattice so cations are next to anions, and not other cations. Vaporisation of ionic materials usually leads to vapours containing discrete ions and some clumps (such as ion-pairs).</p>
A substance with a lattice of covalent bonds, such as diamond (C or C <sub>∞</sub> ) or silica (SiO <sub>2</sub> )	<p>As covalent bonds are directional and strong, these substances tend to have high melting temperatures. Melting the material requires breaking (not just weakening) of the bonds.</p>
Molecular solids	<p>Molecular solids tend to have relatively low melting and boiling temperatures (e.g. nitrogen, methane, carbon dioxide, ammonia etc are in the gaseous state at room temperature), as the bonds between molecules are weak, and the bonds within molecules do not need to be broken for the change of state.</p>
Hydrogen bonded solids	<p>Solids with hydrogen bonding tend to have higher melting and boiling temperatures than other materials with similar size molecules. In ice, water molecules form a lattice with each molecule hydrogen-bonded to four others. This is disrupted on melting, although there is a constant flux of hydrogen bonds being formed and broken in the liquid.</p>

## Dissolving

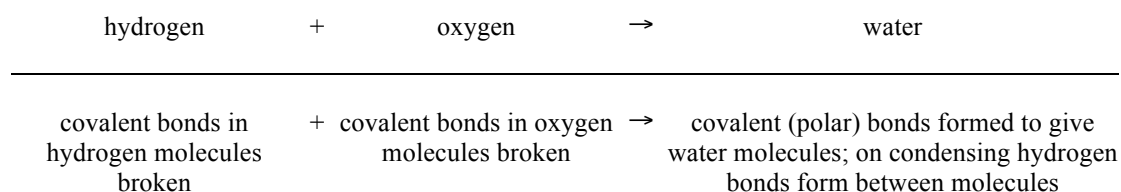
Dissolving of a substance in the solid state involves breaking of bonds in the solid, disruption of bonds in the solvent, and formation of new interactions between solvent and solute.

example	involves	feasible
alloying	Disruption of original metallic lattices and formation of new lattice – however retains delocalised electrons between positive cores.	Often feasible – some mixtures allow a better ‘fitting’ lattice than in pure metals.
salt in oil	Would require breaking strong bonds between ions, but the ions would not bond strongly to non-polar molecules.	No significant dissolving occurs.
salt in water	Requires breaking strong bonds between ions, and disrupting hydrogen bonding in liquid water, but ions often become strongly hydrated as polar water molecules are attracted to the ions.	Some, but not all, ionic solids are very soluble in water.
wax in water	The forces between wax molecules tend to be modest, but hydrogen bonding in water is quite significant, and only weak interactions are formed between water molecules and wax molecules.	No significant dissolving occurs.
wax in oil	The interactions formed between non-polar solvent and non-polar solute molecules are similar in kind and strength.	Mechanical agitation of the solid by collisions from solvent molecules is sufficient to allow a solution to form.
glass in oil or water	Strong covalent bonds in materials such as glass are difficult to break, and would not be compensated for by solvation interactions.	No significant dissolving occurs.

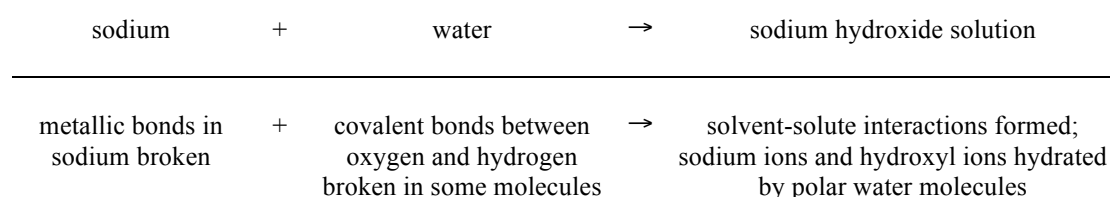
## Bonds and chemical reactions

Chemical reactions seldom occur between substances in an atomic form. Although some reactions of the noble gases have been achieved, these are exceptions. Generally a chemical reaction, such as the various ones described in the chapters of this Handbook, involves both the breaking of bonds in the reactants, and the formation of bonds in the products. Consider a few examples:

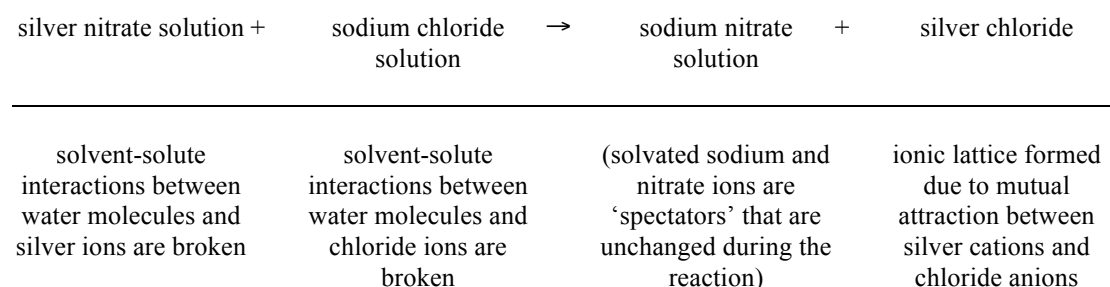
e.g. 1: covalent bonds broken, and formed:



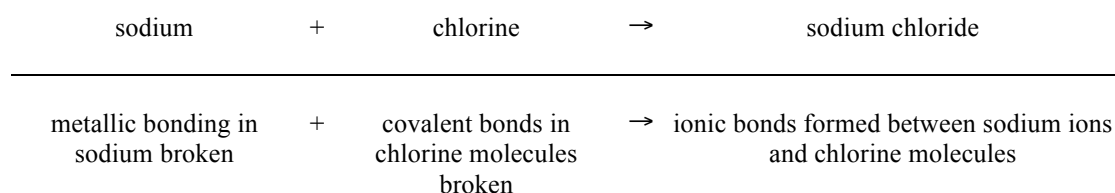
e.g. 2: metallic and covalent bonds broken:



e.g. 3: ionic bonds formed:



e.g. 4: metallic and covalent bonds broken, ionic bonds formed:



Once students have been taught about bond types, it is useful when discussing different reactions in various topics to ask students about the types of bonds broken and formed. This will reinforce learning, help to shift them from thinking of bonding and reactions in 'octet' terms, to electrical interactions, and encourage them to try to visualise what is occurring at the submicroscopic level.

### ***Further reading and resources***

The Royal Society of Chemistry (RSC) publish a range of resources to support chemistry teaching in schools. *Chemical Misconceptions - Prevention, Diagnosis and Cure* (2 Volumes, Taber, K. S., 2002). Includes resources for probing student thinking, and finding out whether students have acquired common misconceptions on a range of topics relevant to this chapter. The probes themselves can also be downloaded from the RSC site:

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

A range of simulations of downloaded from the Chemistry Experiment Simulations and Conceptual Computer Animations page from the Chemical Education Research Group at Iowa State University. These include a simple simulation of a particle model of NaCl dissolving in water, and a simulation showing hydrogen bonding between water molecules in the liquid phase.

<http://www.chem.iastate.edu/group/Greenbowe/sections/projectfolder/simDownload/index4.html>

The simulations available from The Concord Consortium are mostly suitable for more advanced learners, but some of them will be useful for supporting learning with some secondary students. For example, there is a simulation showing how the electron density around two hydrogen atoms are distorted as one is moved towards and then away from the other; and a model showing that the distinction between ionic and covalent bonds reflect the extremes of a bond between two atoms where their electronegativities may be varied by 'sliders':

<http://www.concord.org/activities/subject/chemistry>



