Paper

Undergraduate students' understandings of entropy and Gibbs free energy

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Introduction

For many students, the study of thermodynamics presents problems; it is seen as consisting almost entirely of equations which are not understood and which have to be learned by rote in order to do calculations and to pass examinations. This paper describes part of a study designed to explore students' difficulties in understanding thermodynamics. The paper focuses on students' understanding of entropy and Gibbs free energy.

There have been a large number of studies that explore students' understanding of different science topics at different levels of the education system and which show many mismatches between scientifically accepted concepts and students' conceptions (Pfundt and Duit;[1] Carmichael et al.[2]). The source of students' alternative conceptions lies in how they construct knowledge. When students construct their own meanings they are influenced by their existing (often incorrect) conceptions (Osborne and Wittrock[3]). Knowledge is constructed through a process of interaction between conceptions that already exist in the learner's memory and an outside stimulus; so that when a student sits in a lecture theatre listening to a lecture, she has previously constructed frameworks of conceptions in her memory and recalls these to interpret the new stimulus from the lecture. This process of interaction can cause existing conceptions to be modified or new ones to be created. There are many ways in which this learning process can create unanticipated learning. Two difficulties relevant to this study were, firstly, what students already knew and how knowledge was organised in the memory store and secondly, the kinds of processing required in the interaction between knowledge in that memory store and the new stimulus.

One possible cause of unanticipated learning is a mismatch between students' prior knowledge and the expectations of the teaching staff. If students do not have the requisite prior knowledge, then building new concepts is difficult (White[4] pp.12-21). Alternatively, if students have already constructed understandings and also arranged them in their memory stores in ways that are incompatible with the accepted scientific concepts, new learning is very difficult. Chi et al.[5] argued that knowledge is organised in the memory in different categories. According to these authors, all scientific entities fall into three categories: Matter, Processes and Mental States. If concepts are placed in the incorrect category, then producing conceptual change to result in the accepted scientific understanding is difficult to achieve. There are many examples in learning chemistry where students incorrectly classify processes as things. Some students see the blue colour of copper sulphate as a separate substance, like a pigment in the copper sulphate, and similarly see the sweetness of sugar as being caused by a component that can pass from sugar to water to give water a sweet taste (Sanmarti, Izquierdo and Watson[6]). Students also have difficulties with some chemical reactions because they view heat as a substance taking part in the reaction (Watson, Prieto and Dillon[7]). At a higher level (years 12 and 13), students find ionic bonding difficult to explain because they view bonds as entities rather than interactions between charged particles (Boo and Watson,[8] 2001). It is not surprising that students hold these views, as historically, differentiating substances from properties was one of the key challenges for early chemists. In addition, remnants of these early ideas persist in some of the language used to explain modern chemistry. For example, the properties of classes of organic chemicals are explained by their having certain functional groups. Even the names of some chemicals appear to contain the vestiges of the idea that properties are to be explained by entities rather than interactions. Oxygen, for example, is derived from Greek words meaning 'sour' (oxy) and 'I produce' (gen) because Lavoisier found that the substances he burnt in oxygen produced acids. The problem of placing conceptions into the wrong category was also found to be an issue in the current study.

The second difficulty lies in the cognitive demands that are needed in the processes of interaction between conceptions that already exist in the learner's memory and the new input. It is generally agreed that thermodynamic concepts are abstract. This notion is supported by Dixon and Emery[9] who devised a way of categorizing concepts depending on their level of abstraction; they identified seven such levels. In this scheme entropy was placed on the fourth level (one below enthalpy) while Gibbs free energy occurred on the sixth level. The abstract nature of thermodynamic concepts makes heavy demands on the cognitive processing of students when constructing new meanings. In addition, thermodynamics usually involves the manipulation of two or more variables simultaneously. According to a study by Rozier and Viennot,[10] students treat such systems in a series of sequential steps instead of dealing with the effects of changing all the variables at the same time. Such linear causal reasoning is an example of concrete operational thinking, whereas dealing with the effects of changing all the variables at the same time is an example of formal operational thinking.

The present study took place in England. At the time of this study, successful students in their year 11 examinations could choose to study 'Advanced-level' subjects. They usually chose just three. Although the core Advanced-level syllabus in chemistry contained no reference to entropy or to Gibbs free energy, some A-level syllabuses (e.g. Nuffield, Further Physical Chemistry module of NEAB) did include these concepts. Some students on entry to university could, therefore, be expected to be acquainted with the concepts of entropy and Gibbs free energy. The aim of the present study was to explore students' understanding of these concepts both before and after a first year undergraduate course in chemical thermodynamics, with a view to drawing conclusions about possible improvements in teaching strategies.

There have been a small number of studies of these concepts both at school level and at university. Two studies at school level have been reported: one Scottish syllabus (Certificate for Sixth Year Studies) required equilibrium to be taught from the standpoint of thermodynamics instead of the more usual approach where equilibrium is defined as the dynamic state in which the rates of the forward and back reactions are the same. Johnstone et al.[11] report that students who had been taught this syllabus developed a number of alternative conceptions. These included confusion about thermodynamic reversibility and the failure to appreciate that endothermic reactions could be spontaneous. Boo[12] found that A-level students rarely cited entropy as the driving force of reactions.

The few studies that have taken place in universities show that undergraduate students have a weak understanding of the concepts of entropy and Gibbs free energy (Sozbilir,[13] Pinto,[14] Ribiero et al.,[15] Thomas and Schwenz,[16] Selepe and Bradley[17] and Banerjee[18]). Sozbilir[13] distributed questionnaires to Turkish undergraduate chemistry students in three universities, both prior to and after their course in chemical thermodynamics. Only a small minority of students showed a 'sound understanding' of the concept of entropy and Gibbs free energy. This weak understanding was mirrored in Pinto's[14] study of Catalan undergraduate physics students, who did not connect different aspects of the Second Law with one another and few students used entropy to explain everyday processes. Sozbilir[13] found that students attempted to explain entropy as 'disorder'. However, 'almost all of the respondents defined entropy from the visual disorder point of view, indicating chaos, randomness or instability in some cases'. He found that the term 'disorder' was used to refer to movement, collision of particles and the extent to which things were 'mixed up'. Students did not have clear understanding of enthalpy and energy of a system and seemed to confuse the kinetic energy of a system and entropy. Some also confused enthalpy with Gibbs free energy. Similar findings were found by Ribiero et al[15] and Selepe and Bradley.[17] No student used microstates to explain disorder (Sozbilir[13]).

Another difficulty was understanding the term 'spontaneous'. Few appreciated the thermodynamic meaning of the term and believed that reactions were not spontaneous even when ΔG was less than zero. (Ribeiro et al.,[15] Thomas and Schwenz[16]). Students also confused thermodynamic stability with kinetic stability, believing that a large value for ΔG would cause a fast reaction.

Methodology

The methodology for this study was influenced by the literature survey, which had shown that current research about students' understanding of entropy and Gibbs free energy lacked depth. Previous studies had identified weakness in understanding thermodynamics concepts, as well as some individual alternative conceptions, but the studies failed to give a picture of why misconceptions developed and how they linked to different aspects of students' understanding of chemical thermodynamics. It was, therefore, decided to place a strong emphasis in this study on exploring thoroughly the qualitative understanding of a small number of students, both before and after a taught course of thermodynamics. Such an approach required the use of individual interviews. This type of case-study approach is not intended to generate quantitative data representing all undergraduate students studying chemistry in undergraduate courses in the UK. Rather, it is intended to provide a rich description of the qualities of thinking found amongst the sample. Some quantitative data are presented in this paper, but because of the small sample size these cannot be generalised to represent the numbers of students who would be expected to hold such views in the total year group. The data do, however, show where students had difficulties in understanding and where their common misconceptions were. Previous studies have shown (e.g. Marton[19]) that there is some stability in the kinds of conceptions found in groups of students, and it is therefore expected that other groups of students would exhibit conceptions of similar qualities. A reader can generalise from this small sample in so far as he or she recognises common features between the case described here and his or her own experience.

The sample

The sample was drawn from a first year undergraduate cohort of 100 students attending a university chemistry department in England, which has a good reputation for Chemistry (as measured by external evaluations of research and teaching quality). A sample of 20 was chosen randomly. The grades achieved by the sample and the whole cohort in their pre-university examinations (the English Advanced-level examinations taken at age 17-18) were similar: i.e. 22.4 ± 2·8 and 19·3 ± 2·9 respectively, as measured using the numerical values for A-level grades across all subjects. Students were interviewed twice, once before a lecture course in chemical thermodynamics and once after the course. Because of difficulties in tracking down some students and the imminence of examinations, only 16 of the original 20 appeared for the second interview. The results, which follow, refer to the 16 students present at both interviews. All the interviews were tape-recorded and transcribed in full.

The teaching

The lecture course consisted of 13 one-hour lectures at the rate of two per week. Students were also expected to attend weekly examples classes for which problems relevant to the course were set in advance. The researcher observed all the lectures and examples classes and took field notes on the content of the lectures and the students' behaviour (such as whether or not they were attentive or restless and when they made notes). A transparency was used to summarize the main course content; this was displayed at the beginning of each lecture, and progress discussed. It was noted that all equations that arose during the lectures were written up on the blackboard, interspersed with very occasional written summaries of more important points. It was noticeable that students only copied down material from the blackboard; only rarely did they make notes on any spoken material. In the examples classes, the students worked in small groups (2-3 students) solving problems presented in advance. The three or four staff members present helped students when required. All the problems set were calculations based on the current week's lectures.

The lecture course included the following main concepts:
Energy levels, calculation of the energy of a collection of molecules; internal energy, U; first law of thermodynamics, heat, work; enthalpy, H; different distribution of molecules with the same energy; entropy, S; calculation of entropy for occupations of different energy levels; effect of changing temperature on energy level spacings; distribution of particles over lattice sites; standard entropies; heat capacity $C = q/\Delta T$, measurement of C_p and C_v; second law of thermodynamics, equilibrium, Gibbs energy, G; measurement of equilibrium constants and calculation from Gibbs energy; calculation of yields of reactions; ionic product of water, pH , weak and strong acids.

The statistical approach to explaining entropy was built on an earlier course in quantum mechanics in which the idea of energy levels had been explored. The concepts of entropy and Gibbs free energy, a subset of all the concepts in the course, have been broken down and form the list in Table 1.

In order to place this course in a wider context, eight other university chemistry departments were approached in order to ascertain the general thrust of their thermodynamics teaching. Analysis of the syllabuses of these eight departments revealed a very similar approach both in terms of the content of the courses and the teaching approaches. All the courses covered the same concepts, although some also dealt with more advanced concepts such as chemical potential. Courses emphasised the mathematical aspects of thermodynamics and were delivered by means of formal lectures, supported in some cases by examples classes.

The interviews

The interviews focused on three chemical reactions, which should have been familiar to the students. The reactions were:

- the neutralization reaction between hydrochloric acid and sodium hydroxide solutions, both at concentration 2 mol dm^{-3}. The reaction was exothermic and the only visible sign of reaction was the increase in temperature registered on the thermometer;
- the reaction between magnesium (ribbon) and hydrochloric acid (2 mol dm^{-3}). Again, this was exothermic but visible changes were also observed, namely the effervescence caused by the evolution of hydrogen and the ‘disappearance’ of the magnesium;
- the dissolution of ammonium chloride in water. This was included as the only endothermic reaction. Students, whose understanding depends on the idea that a reaction can only happen if there is a fall in energy level from reactants to products, find it impossible to explain why this reaction should occur.

After a preliminary discussion about each reaction, which included writing the equation for the reaction and describing any observations they had made, students were asked to respond in thermodynamic terms to questions about the reactions, e.g. ‘What happened in this reaction to cause the temperature change?’ and ‘Why did the reaction happen?’ Each interview consisted of two parts. In the first part, the chemical reactions were used to focus students’ ideas. During this, open questions were asked so that the students could decide the terms in which to frame a response; students referred to the concepts of entropy or Gibbs free energy only if these seemed relevant to them. Supplementary questions were asked so as to ascertain what was meant by each response. During the second part of each interview, students were asked directly what they understood by the terms ‘entropy’ and ‘Gibbs free energy’. It was in this final part of the interviews that students who were unsure about the concepts produced most alternative conceptions.

The first step in the analysis consisted of developing a set of statements to represent scientifically accepted statements about the concepts of entropy and Gibbs free energy (Table 1). These lists were closely related to the content of the lecture course and were a subset of all the concepts covered in the course. Each student transcript was carefully studied and marked up so that each student statement was related to the statement to which it referred; if it was reasonably close to the accepted scientific view, then that statement was marked as ‘correct’. If, however, the student’s ideas did not correspond with the scientific view, it was identified as an alternative

Table 1

Scientifically accepted statements about the concepts of entropy and Gibbs free energy

1. Entropy gives a measure of the number of ways in which energy is distributed among energy levels within and between particles (or the number of microstates).
2. The more widely spread the energy quanta among the various energy levels, the more probable the state and the higher the entropy.
3. For a chemical reaction to be possible, the total entropy change ($\Delta S_{system} + \Delta S_{surroundings}$) must be positive or not negative.
4. A spontaneous reaction is one that is thermodynamically feasible; that is, one for which the total entropy change ($\Delta S_{system} + \Delta S_{surroundings}$) is positive. The reaction may not occur because of kinetic barriers.
5. If energy transferred from an exothermic reaction increases the temperature of the surroundings, the entropy of the surroundings is increased. [$\Delta S_{surroundings}$ can be calculated from the expression: $\Delta S_{surroundings} = -\Delta H/T$].
6. The entropy of a system can increase when:
   (a) it gains energy;
   (b) a change of state occurs from solid to liquid to gas;
   (c) mixing of substances occurs;
   (d) the number of particles increases during a reaction.
7. In order to predict whether a reaction is possible it is necessary to know either
   (i) the total entropy change which must be positive or zero, i.e. not negative, or
   (ii) the free energy which must be negative.
8. Gibbs free energy, G, is the maximum amount of energy that is available from a chemical reaction to do useful work (other than pV changes which are not available to do useful work). Alternatively: Gibbs free energy is the energy available to do useful work after any energy transfers have taken place which ensure that the total entropy change ($\Delta S_{system} + \Delta S_{surroundings}$) is not negative.
9. Gibbs free energy must be negative for a reaction to be possible.

conception of that statement. A further step in the analysis consisted in identifying commonalities between alternative conceptions related to the different statements in Table 1.

Results

An overview of the understandings of students before and after a lecture course is given below, followed by a detailed examination of some of the students' responses. Quotations from interviews include first the student number, and then a roman numeral, I, to indicate the interview was before, or II, to indicate that it was after, the lecture course. As the interview transcripts were read, a note was made whenever a student made a statement that showed he or she had understood one of the statements in Table 1 or had an alternative conception. During the course of an interview, a student often changed his or her mind, thus giving rise to the recording of more than one conception; sometimes all the conceptions shown by a student were alternative ones, sometimes one was correct and the others alternative.

Table 2 shows the numbers of correct and alternative conceptions that were recorded for each statement both before and after the lecture course. Sometimes the total number of conceptions recorded is greater than 16 (the number of students) because individual students used more than one explanation for a particular concept.

Before the lecture course five students had heard of the concept of entropy from their school chemistry courses. Their knowledge of some aspects of entropy was classified as being scientifically acceptable, but they did not produce satisfactory definitions of entropy in terms of microstates (or even of ways of arranging particles). They knew enough to say, with some degree of confidence, that total entropy change had to be positive when explaining why the reactions happened (statement 3). Two of these students suggested entropy as randomness or disorder. However, they did not explain what it was that was random or disordered. Some other students claimed to have heard of the term 'entropy', but had no understanding of its meaning. The large number of alternative conceptions in the interviews before the lecture course was in response to the direct question in the second part of the interview asking students what they understood by the term 'entropy'. Often it was suggested that entropy was another 'form of energy', and entropy was occasionally confused with enthalpy as the words are similar. The prevalence of the 'forms of energy' framework in students' thinking has been discussed in relation to students' understanding of enthalpy in an earlier paper (Carson and Watson[20])

The few who had heard of Gibbs free energy from their school chemistry courses did not explain what it meant, but were aware that the change in free energy had to be negative (statement 9) for a reaction to be possible.

After the lecture course there was a much greater awareness of the term 'entropy', and a large increase in the number of students with scientifically acceptable concepts. In particular, more students had learned that for a chemical reaction to be possible the total entropy change

Table 2

Numbers of students with scientifically accepted ideas or alternative conceptions about entropy and Gibbs free energy

Proposition	Before the lecture course		After the lecture course	
	Number of scientifically accepted ideas	Number of alternative conceptions	Number of scientifically accepted ideas	Number of alternative conceptions
1	0	12	6	11
2	0	1	7	4
3	4	3	11	7
4	2	0	3	8
5	3	2	6	7
6	2	2	12	9
7	4	10	5	8
8	0	6	0	5
9	4	2	10	6
Total	19	38	60	65

must be positive (statement 3) or that the change in Gibbs free energy must be negative (statement 9), and had learned the ways in which entropy of a system can increase (statement 6). Gibbs free energy was a concept which students found particularly obscure even after the lecture course. Though the term had become familiar, none explained its meaning (statement 8) but the number who now knew that it had to be negative for a reaction to be possible had increased (statement 9). However, alongside the increase in scientifically acceptable explanations, there was also a large increase in the number (from 36 to 65) of alternative conceptions.

During the lecture course, entropy was explained in terms of the distribution of energy among available energy levels. Therefore, some explanations of the nature of entropy in terms of energy levels or microstates would be expected from the students. However, very few of them seemed to have taken these ideas on board. As can be seen from Table 2, after the lecture course six students produced explanations in which energy levels were mentioned (statement 1) and attempts were made to clarify their explanations using diagrams of energy levels, e.g.

> *Student 1 (II) line 26 and 30: "It [disorder] relates back to the distribution of molecules in microstates, how the spread of microstates; it's one way of positioning the molecules over the energy levels."*

Such student responses appear as 'correct' in the Table 2. As can be seen from this example, even when familiar with the idea of microstates, students found it difficult to express their understanding in words. Most students still tended to refer to ideas of randomness or disorder and were no clearer on what these terms meant than before the lecture course. It was noticeable, however, that the ability to apply their ideas about entropy to chemical reactions had improved considerably.

Even when students appeared to have grasped the idea that since a reaction had happened the total entropy must have increased, some showed their understanding was flawed because they ascribed the negative change in enthalpy during the reaction to the increase in entropy:

> *Student 17 (II) line 10: Well there's a spontaneous reaction so it's been a change in enthalpy heat's been evolved which is because the entropy is favourable so well spontaneous reaction proceeds.*

Also, most students did not differentiate clearly between the system and the surroundings (statement 5), making it difficult to determine whether they understood that total entropy had increased.

Students found it easiest to relate entropy changes to changes of state, and some of the simpler ideas about entropy (that the entropy of substances increases in the sequence solid to liquid to gas) were well known and were used successfully in explanations. Such explanations can be accounted for if one realises that to these students the particles involved were concrete entities, and students' mental models involved images of the increasing randomness in the arrangement of particles in going from solid to liquid to gas.

After the lecture course students recognized Gibbs free energy as an entity whose relationships could be manipulated in calculations in order to predict whether or not reactions could happen. However, students showed no understanding of the concept. The only definition given in the lecture course was the relationship: $\Delta G = \Delta H - T\Delta S$. It is, therefore, not surprising that there were few alternative conceptions identified; students did not find the relationship intelligible and they were satisfied with their current framework in which Gibbs free energy was seen as another 'form of energy'.

Analysis of the transcripts from both sets of interviews revealed several ideas or misunderstandings that seemed to underlie students' alternative conceptions of entropy and Gibbs free energy:

forms of energy explanations,
explanations of entropy in terms of disorder or randomness,
entropy changes explained solely in terms of change of state,
confusion between system and surroundings.

Forms of energy explanations

In these explanations students showed some evidence of awareness that energy was in some way involved but did not clearly differentiate between enthalpy and entropy or between Gibbs free energy and entropy, seeing them all as simply 'forms of energy':

> *S1 (I) line 289: …like it's [entropy] another name for enthalpy.*
> *S2 (II) line 150: … gases have a much higher energy a much higher chaotic energy the entropy it is much greater.*

In these statements students are using entropy, enthalpy and chaotic energy as terms to simply label what they see as different forms of energy. These statements were typical of many instances of students using the underlying alternative framework of 'forms of energy' to try to make sense of thermodynamic terms.

Explanations of entropy in terms of randomness or disorder

Students using this idea talked about randomness or disorder but failed to explain what these terms meant. Such students made no mention of microstates nor of energy levels and simply explained entropy as:

> *Student 6 (I): randomness or disorder of everything.*

Entropy changes explained solely in terms of change of state

Some students explained entropy changes solely in terms of changes of state. They seemed to recall that there was an increase in entropy in changing from a solid to a liquid to a gas, but did not give the underlying explanation in terms of increasing the ways in which energy could be distributed:

> *Student 2 (II) line 106: You've obviously got an increase in the amount of entropy going on . . . the solid then changing the phase . . .and just changing the phase will increase the entropy of the system and again you've also got a change of state because you've got a gas given off at the end of the process where you didn't have any to begin with so you've also got more entropy in your system because there's phase changes going on there as well. If you cross phase boundaries the entropy is increased.*

Students' use of change of state to explain entropy appears to be linked to a view of entropy as randomness of movement of particles.

Confusion between the system and surroundings

There was frequent confusion about *system* and *surroundings*, often with the surroundings being ignored. This resulted in a lack of understanding of the effects of the transfer of energy to and from the surroundings. In the example below the argument is made from the standpoint of the system alone, neglecting any changes in the entropy of the surroundings:

> *S19 (II) line 50: If it's thermodynamically favourable a reaction will go spontaneously so generally speaking the entropy change is negligible compared with if the enthalpy change is favourable i.e. if it's large and negative enthalpy change the reaction will occur.*

Discussion

This study confirms earlier ones in showing that undergraduate chemistry students have already formed some misconceptions and more general alternative frameworks related to chemical thermodynamics before they begin their university studies, and that these have a significant impact on their understanding. Previous studies have identified confusion in the use of various thermodynamic terms like enthalpy, energy, entropy and kinetic energy (Sozbilir,[13] Ribiero et al.,[15] Selepe and Bradley[17]). This study goes further in that it identifies a particularly persistent alternative framework that interferes with students' understanding of new concepts in thermodynamics, i.e. 'forms of energy'. This framework, which is used in lower secondary schools, views energy as a quasi-material substance that can be transferred between entities and can take different forms. The students perceive energy as being in the category of 'matter' (Chi et al.[5]). The essence of thermodynamics, however, is a study of interactions. Terms like entropy and Gibbs free energy cannot be understood as isolated entities that can be transformed into one another. If students are to use these concepts to make predictions about whether reactions can occur, they need to understand them in the context of chemical 'processes' (Chi et al.[5]). This change in the way of thinking about chemical reactions is very difficult to achieve and involves radical conceptual change (Chi[21]). It also requires the use of more complex thinking (Rozier and Viennot[10]).

Another example of students' difficulties in coming to terms with complex abstract ideas is their explanations relating to the use of the concept of entropy. Entropy was described in vague terms such as chaos or randomness, often without specifying what was chaotic or random. When students tried to be more specific, they related entropy changes to changes of state rather than to distribution of energy in microstates. Sozbilir[13] reported a similar finding. In that case students seemed to be using a concrete analogical model related to ideas about kinetic theory that they had learned at school. Harrison and Treagust[22] have shown, however, that students find it difficult to replace such concrete analogic models with abstract ones.

The results above reveal a mismatch between the learning needs of students and the contents and the approach of the thermodynamics course that they studied. The lecture course, examples classes and examination all emphasised numerical calculations using thermodynamic equations. There were no opportunities to elicit students' qualitative explanations to find out what they understood before they started the course and little opportunity to develop their qualitative understanding during the course. Instead, students learnt to manipulate symbols without understanding the concepts that they represented. By presenting thermodynamic definitions only in terms of mathematical relationships, for example by defining Gibbs free energy solely as the relationship: G = H – TS, students are allowed to ignore the intrinsic meaning of the expression while they concentrate on using it to perform calculations. Students do not automatically invest such expressions with all the meaning that experts in thermodynamics bring to bear and it is unreasonable to expect them to do so. It is necessary, therefore, that thermodynamic entities such as entropy and Gibbs free energy are defined qualitatively before the introduction of the mathematical expression.

Implications for teaching

The results reveal the strong influence of students' prior learning on the development of thermodynamics concepts. Effective teaching needs to take into account conceptions held by students before they start the course and those alternative conceptions developed during the course. At the beginning of a course, a seminar might be held in which students would be invited to write or talk about the thermodynamic implications of some familiar chemical reactions. This would reveal the major alternative conceptions held by that group of students. Also, it would be profitable to challenge students with some already known alternative conceptions, asking them to support or deny these conceptions and explain their answers.

The main emphasis of chemical thermodynamic courses seems to be on the quantitative and mathematical aspects of the subject. One of the difficulties with this approach is that students learn by rote the facility to carry out calculations – correct answers gain good marks. They have no incentive to construct proper meanings for the ideas involved in the calculations. Concepts such as entropy and Gibbs free energy need to be described qualitatively. Because students do not read into mathematical relationships all the meaning that an equation carries, teachers need to provide students with qualitative explanations of such thermodynamic relationships.

Another influence on student learning is the style of examination questions. Questions need to be of a kind that require students to demonstrate an understanding of the concepts involved. The setting only of numerical calculations serves to emphasise that competence in manipulating equations is all that is needed to learn thermodynamics. Entwistle and Entwistle[23] found that, even when the lecturer had admirable aims in terms of conceptual understanding, unless examination papers reflected this, students were strongly influenced into rote learning.

It is necessary, too, to discuss directly the limitations of alternative frameworks such as 'forms of energy' by pointing out that, while as a model it was satisfactory in school, it is now necessary for students to realise that, for example, entropy is not just another 'form of energy'. They need to accept and to think of entropy as entropy. What is important, in general, is that teachers should address explicitly some of the mental models held by students and compare these directly with new models being presented to them in the thermodynamics course.

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