

Questions that science teachers find difficult (II)

Preguntas consideradas difíciles por los profesores de ciencias (II)

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Abstract

This is the second of a short series of articles in which the author discusses some more of the questions that he has found difficult - and still finds problematic - in teaching science. Difficulties arise from a number of perspectives including (a) the teacher's own understanding of the science and (b) making accessible to pupils/students an answer or explanation that is appropriate for them. This paper focuses upon three further questions relating to 'simple' everyday situations that are normally explained in terms of the kinetic theory of matter. The third question - looking at the difference between chemical and physical changes - is becoming increasingly problematic for the author.

Key words: difficult concepts, science, chemistry, physics

Resumen

Este es el segundo corto artículo a donde el autor quiere discutir otros conceptos que él encontró difíciles o problemáticos en Educación en Ciencias. Estas dificultades aparecen desde diferentes perspectivas : a) el propio entendimiento del profesor de Ciencias; b) desde el punto de vista de acceder el entendimiento de los estudiantes a través de la explicación correspondiente. El enfoque de este trabajo está en tres situaciones de la vida cotidiana lo cual normalmente se explica en términos de la teoría cinética. La tercer situación sobre las diferencias de los fenómenos químicos y físicos, también ésta se convierte en problemática para el autor.

Palabras Clave : conceptos difíciles, ciencias, química, física.

Introduction

In the previous paper (Goodwin (2002a)) I expressed the hope that readers would respond to the questions either with criticism or comment, answers, or upon the issues raised. To date the response has been complete silence and it is difficult to draw conclusions from this. Perhaps there are no readers 'out there'? Perhaps you think the questions are too simple, too hard or simply irrelevant? Perhaps the issues do not arise in your teaching? Please let us have *your* views - your *questions* either directly or via the editor.

The first paper reached question 3 and thus this paper begins with question 4.

Difficult Question 4

Why does the temperature of the mixture fall when salt is stirred into melting ice?

A mixture of water, ice and salt (sodium chloride) is often called 'freezing mixture' and has long been used in homes and in laboratories as a simple way of producing temperatures well below 0°C. Well before the invention of refrigeration powered by electricity, mixtures of ice and salt were used to reach temperatures low enough to make ice cream. Fahrenheit used the lowest temperature that he could achieve by mixing ice and salt to fix the zero point on his scale of temperature. On the Fahrenheit scale the freezing point of water is designated +32 °. (The lowest possible temperature that could be achieved by mixing ice and salt is -21.2 ° C or almost -6.2 ° F. This is called the 'eutectic' temperature.)

My answer to myself: Explaining why the temperature falls is not easy. It is 'correct' to state that dissolving salt (or any other substance) in water reduced the temperature at which the solution will begin to freeze. However, this does not really explain what is happening. Nonetheless it is important that teachers and pupils are convinced that this *is* the case before any explanation can make sense.

There is a convenient, although *incorrect*, explanation that is based on an alternative conception that is attractive to both teachers and students. It seems to make sense that the separation of the positive sodium ions from the negative chloride ions as the crystals dissolve in water requires a lot of energy, and that it is to supply this energy that the temperature falls. However, there is no marked fall in temperature when salt dissolves in water so the explanation cannot be valid. The energy given out when water molecules hydrate the sodium and chloride ions is almost the same as the energy required to separate them, so the energy change is close to zero when salt dissolves in water. (Goodwin, 2002b).

An acceptable explanation seems to be possible only when we have some understanding of the criterion

for equilibrium between a solution and the solid solvent. It is the temperature of this equilibrium that is the freezing or melting temperature. At this temperature and at equilibrium the *saturated vapour pressure (SVP) of water over the solution and the solid ice are the same.*

When salt is added to melting ice the salt dissolves in the water, *not in the ice*, and reduces its SVP. Thus to attain equilibrium some of the ice melts as water molecules move from the state with higher SVP to that with the lower. The necessary potential energy is gained from the conversion of kinetic energy of molecules in the system, which results in a fall in temperature. Equilibrium is re-established when at the lower temperature the SVP of ice is again equal to that of the (now more dilute) salt solution.

Difficult Question 5: How is it that endothermic processes can occur spontaneously?

At an elementary stage of learning examples of these are not abundant, but they are not so uncommon that they can be ignored.

Examples at room temperature:

- a. Dissolving of most solid solutes in water (see question 4 above)
- b. The contraction of stretched rubber
- c. Evaporation of liquids/solids (see question 2 in previous paper)

Mostly these would *not* be considered as *chemical reactions*. Nonetheless they give a lie to the idea that systems can only change spontaneously if they give out energy in the process.

My answer to myself: The issue here is whether, and when, it is appropriate to introduce the idea of ‘entropy’ to students and whether I understand it sufficiently myself to be convinced by the idea in qualitative terms.

It seems to me that it is reasonable to suggest that a system made up of a huge number of small particles in constant random motion is much more likely to be mixed up rather than ordered. Entropy, as a measure of ‘mixed-up-ness’ or ‘disorder’ or ‘the number of ways that the energy of the system can be distributed among the particles’ is not easy to quantify. However, if there is no large energy barrier to prevent mixing it seems reasonable to suppose that a system will become ‘as mixed up as possible’. Also, that the higher the temperature, the more important the tendency towards higher entropy (mixed-up-ness) becomes.

In a pure substance the solid state is more ordered than the liquid. The gas stage is most disordered. It makes sense that for all(?) substances solids are the low temperature form and that liquids and gases form at higher temperatures. (SVP increases with temperature, although the size of the 'energy barrier' also determines when the melting and boiling points occur.)

$$\Delta G = \Delta H - T\Delta S$$

I fear this equation is often learned and applied without any qualitative feel for what it means!

Difficult Question 6: What is the difference between chemical and physical changes?

I used to think this was the easiest possible question and recall learning the following table in the very early stages of science education at about the age of 11 years.

Physical Change	Chemical Change
a. No new substance formed b. Energy change usually small c. Usually easily reversed	a. New substance(s) formed b. Energy change usually large c. Often more difficult to reverse

Table 1: Chemical and Physical Changes

Fortunately the only examples given (or asked in examinations) were fairly straightforward.

Physical: e.g. boiling liquid water to give steam (no new substance, easily reversed), small easy change); melting ice; dissolving salt/sugar in water; melting wax; breaking glass (although difficult to reverse).

Chemical: e.g. burning carbon/magnesium/paper/coal; heating potassium nitrate or mercury II oxide; reacting acids with metals, metal oxides or carbonates.

The ideas were quite useful, especially the criterion as to whether or not a new substance was formed. However, as I have learned more about chemistry, many of the physical changes I once thought were unproblematic have become much more uncertain.

Initially I thought that dissolving 'things' in water would be always physical changes. It soon became clear that some substances (e.g. metallic sodium, potassium, calcium) react with it chemically. Some time later the solution of hydrogen chloride or concentrated sulphuric acid in water were shown to provide solutions with very different chemical properties than the water and the acid separately.* I am now not sure whether hydrated sodium and chloride ions are chemically distinct from the anhydrous ions. If they are chemically distinct, then dissolving salt in water is a chemical change! (Hydrated Copper II ions are blue whereas the anhydrous ion shows no colour – anhydrous copper II sulfate is white - so presumably in this case there is a chemical distinction.)

Further than this: hydrogen bonds are broken when liquid water evaporates. If the hydrogen bond is a 'chemical bond' then surely liquid water molecules are in a different chemical form than in the vapour state? When glass is broken surely *chemical bonds* are broken - although is a new substance formed?

What is a substance? When is an attraction between atoms a chemical bond? I thought I knew the answer forty years ago - now I'm not so sure.

(* I was surprised, for instance, to find that magnesium ribbon does not react with cold concentrated sulphuric acid, although the reaction with dilute acid is extremely vigorous! If you try this, please take great care.)

Bibliography

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