Module 3: Water

Some teaching points

1. Neutral species equations or full formulae equations?

Conquering Chemistry uses the term *neutral species equation* for equations such as Equation 8.4 on page 206 whereas the syllabus document (Section 9.5.4) uses the term *full formulae equation* for such equations. Neither term is completely satisfactory, so teachers and students should feel free to use whichever term they prefer. The important thing is to recognise that there are three types of equations for ionic reactions, complete ionic, net ionic and the one that writes the ions as parts of neutral compounds.

2. Introduction to equilibrium

If equilibrium is introduced before concentration (as is done in the syllabus), care is needed not to use concentration in explaining equilibrium! The only points about equilibrium that need to be introduced here are
(a)   equilibrium reactions do not go to completion, and
(b)   the equilibrium is a dynamic one (both forward and reverse reactions occurring but at equal rates)
 But remember rate of reaction has not been formally introduced at this stage!

Le Chatelier's principle need not be introduced until Module 1 of the HSC course.

3. Different measures of concentration

The various measures of concentration listed on pages 212–3 seem to be required by the syllabus, though they were not in the old syllabus. The reason for deferring discussion of ppm for gases till later is that it needs the molar volume of gases (to interconnect molecules per million molecules and volumes per million volumes) which is not introduced until Module 2 of the HSC course.

4. Exothermic reactions get hotter??

Some students have difficulty with the idea that when an exothermic reaction is performed in a test tube the test tube gets hot: they see this as the test tube and its contents gaining heat not losing it and so shouldn't the reaction be endothermic?. Hence it is necessary to be very careful in describing exo- and endothermic reactions.

In exothermic reactions there is a loss of energy (chemical energy) by the reactants as they form products and this energy is given out as heat and this heat warms up the test tube and its contents. The process is chemical energy being converted to heat energy. Similarly with endothermic reactions: as they occur the reactants need to take in energy (to convert to chemical energy) to store in the products. Generally the only place that the reaction can get the required energy is from the heat in the test tube and its contents, so the tube and contents get cold: heat energy is being taken out of the substances and converted to chemical energy stored in the products of the reaction.

It is perhaps unfortunate that enthalpy (next module, page 276) is often described as 'heat content' (but not in CCPC 4th edition) rather than as the total energy contained in a substance. Despite how enthalpy is defined, the *change in enthalpy* is correctly defined as heat absorbed at constant pressure. The

$H_{\text{reactants}}$ and H_{products} in Figure 10.1, page 276, are really chemical energies rather than heat contents. When there is an increase in chemical energy, ΔH positive, the energy is obtained by extracting heat from the substances and so the reaction mixture gets cold. And when there is a decrease in chemical energy, ΔH negative, the released chemical energy appears as heat and so the reaction mixture gets hot.

5. $\Delta H = - m\, C\, \Delta T$?

The syllabus introduces this equation in Section 8.4.5 but does not state what ΔH stands for. Generally in chemistry texts ΔH means *change in enthalpy*, but enthalpy is not introduced until Module 4. The quantity $m\, C\, \Delta T$ is a measure of the heat absorbed or the heat released when a body or substance undergoes a change in temperature, ΔT. If we define ΔT as the increase in temperature as the sample goes from its initial state to its final state, then $m\, C\, \Delta T$ will be the heat absorbed by the sample. Chemistry books generally use q for the heat absorbed. So the correct equation is

$$q = m\, C\, \Delta T \qquad\qquad\qquad ...(5.1)$$

This is Equation 8.11 on page 223. If the temperature of our sample increases (ΔT positive), then q is positive; that is heat is absorbed by our sample. If the temperature of our sample decreases (ΔT negative) then q is negative. If a negative amount of heat is absorbed, that must mean that heat is *released*.

How does ΔH get into this equation? As we shall see in Chapter 10, ΔH is the symbol used for the change in enthalpy for a chemical reaction (or other process such as the dissolution of a substance in a solvent). ΔH is defined as the heat *absorbed* when the change occurs at constant pressure. As we shall also see in Chapter 10 we often measure ΔH by measuring the change in temperature in a quantity of water or solution in which the reaction (or dissolution) being examined has occurred. For example we determine the heat of solution of a solid in water by measuring the change in temperature in the water when a known mass of solid is dissolved in the water. Suppose the temperature increases as the solid dissolves. We calculate the amount of heat *absorbed* by the water from the above equation. This heat absorbed by the water must be the amount of heat *released* by the dissolution. So the amount of heat absorbed by the dissolution (or reaction) must be minus this quantity. That is:

Amount of heat absorbed by the dissolution or reaction $= - m\, C\, \Delta T$

If the process occurred at constant pressure (that is in a container open to the atmosphere), then the heat absorbed is ΔH; that is

$$\Delta H = - m\, C\, \Delta T \qquad\qquad\qquad ...(5.2)$$

ΔH is the enthalpy change for the process (reaction or dissolution) occurring, while $m\, C\, \Delta T$ is the heat absorbed by the substance receiving the heat from the process. Generally ΔH is quoted per mole of specified substance. What we have so far is the enthalpy change for the amount of substance dissolved or reacted in our experiment so we need to convert it to an amount per mole.

If the experiment was not performed at constant temperature, for example if it was performed in a closed container and involved production or loss of a gas, then the heat released by the reaction (that is the heat absorbed by the water or solution) would not be ΔH. Hence Equation 5.2 applies only under very restricted conditions. On the other hand Equation 5.1 here (or 8.11 in the book) is much more widely applicable; it applies whenever a sample undergoes a temperature change.

Example 13 illustrates. It may be clearer if after 'From Equation 8.11' it had read 'heat released by dissolution of NaOH = heat absorbed by the solution = 211.2 x 4.2 x (31.4 – 19.2)'