What is the relationship between a solvent and the substances that dissolve in it?

In this activity you investigate the solubilities of four substances in three different solvents. You can draw some general conclusions about solubility from the data you collect, and explain your observations in terms of chemical ideas you have learned during the course. You may need to revise ideas about chemical bonding (Chemical Ideas 3.1 and 5.1), intermolecular forces (Chemical Ideas 5.3, 5.4 and 5.6), and the enthalpy and entropy changes during solution (Chemical Ideas 4.5).

Requirements

- the following solids, in powdered form in stoppered specimen tubes (about 2 g of each):
 - anhydrous sodium chloride
 - anhydrous calcium chloride
 - glucose (or sucrose)
- grated candle wax (2g)
- test-tubes (3) and rack
- hexane (12 cm^3)
- propanone (12 cm³)
- distilled water (12 cm³)

CARE Hexane and propanone are highly flammable liquids. Keep bottles stoppered when not in use and well away from naked flames. Avoid skin contact and do not breathe the vapours. Return residues to the correct residues bottle. Do not pour them down the sink.



What you do.

1 Test the solubilities of the solids in the different solvents using the method described in Activity M1.3 of From Minerals to Elements. (CARE Hexane and propanone are highly flammable. Extinguish all flames before using these liquids. Avoid skin contact and do not breathe vapours. Return residues to the correct residue bottle. Do not pour them down the sink.)

It may be best for each pair of students to investigate the solubility of one solid in the range of solvents, and then join with other pairs to produce a full table of results.

2 Use the information contained in your results table to work through the points which follow.

Looking at your results

- a For each of the seven substances used (four solids and three solvents), decide whether its bonding is ionic or covalent. If it is covalent, are the molecules polar or non-polar?
- **b** What general pattern of solubility do you observe from your results?

To interpret your results fully you should read about energy changes in solution in Chemical Ideas 4.5. Entropy changes usually favour solution but the process is often prevented by a large endothermic enthalpy change of solution.

We can find out whether dissolving is likely to be accompanied by a large endothermic enthalpy change by considering the interactions between the particles in the separate solids and liquids, and between the particles in the solutions which might form from them. For example, if the attractive forces between particles in the pure solid and in the pure solvent are stronger than those between particles in solution, the enthalpy change of solution is likely to be endothermic.



WHAT IS THE RELATIONSHIP BETWEEN A SOLVENT AND THE SUBSTANCES THAT DISSOLVE IN IT? 01.1

 ${\bf c}$ $\,$ Fill in Table 1, and use it to explain why solutions form or do not form in each case.

Solid	Principal interactions in solid	Solvent	Principal interactions in solvent	Principal interactions in solution	Enthalpy change which would accompany the formation of a solution as a result of these changes in interactions	Is the solid likely to dissolve?
NaCl		hexane				
		propanone				
		water				
CaCl ₂		hexane				
		propanone				
		water				
glucose		hexane				
		propanone				
		water				
wax		hexane				
		propanone				
		water				

Table 1 Changes in interactions between particles when a solid dissolves



What changes occur when an ionic solid dissolves? This activity looks in more detail at the changes which occur when ionic solids dissolve. It provides evidence to support and test the explanations you put forward in Table 1 in Activity O1.1. It may be best to share the work so that each student or pair of students chooses one solid to study in both parts of the activity.

calcium chloride

iron(III) chloride

CARE Eye protection must be worn.

Requirements.

- the following solids, in powdered form in stoppered specimen tubes:
 - anhydrous sodium chloride (12g)
 - anhydrous calcium chloride (23g)
- anhydrous iron(III) chloride (33g)
- distilled water (300 cm³)
- polystyrene cups (or insulated beakers)
- 0–110 °C thermometer
- 50 cm³ measuring cylinder
- burettes (2)
- weighing bottles with stoppers
- rubber bungs to fit burettes
- small funnel
- access to balance

What you do.

Note Calcium chloride and iron(III) chloride each take in water from the air. It is important to use *fresh anhydrous* samples and to work quickly.

Part 1: Enthalpy changes

- 1 Place 50 cm^3 of distilled water in a polystyrene cup and record its temperature.
- **2** Weigh out enough solid to produce a solution of concentration 2.0 mol dm⁻³.
- **3** Add the solid to the water. *Carefully* stir the mixture with the thermometer and record the highest or lowest temperature of the solution.
- 4 Calculate the temperature change and from it the molar enthalpy change of solution ($\Delta H_{solution}$) (i.e. the enthalpy change when 1 mole of solid dissolves to form a 2.0 mol dm⁻³ solution).
- 5 Repeat the procedure in steps 1–4 for the other two solids.
- 6 Draw up a table to show your results.

Part 2: Volume changes

This part of the activity uses a burette to make accurate measurements of the volume changes which accompany the formation of a solution. Burettes contain an ungraduated section and you have to be careful when filling the burette to make sure it contains only 50 cm^3 of water. One way of doing this is to add exactly 50 cm^3 of water from another burette (or a pipette) to your empty burette.

- 7 Weigh out enough solid to make 50 cm^3 of a 2.0 mol dm⁻³ solution.
- 8 Place 50 cm³ of water into the burette, then pour in a little of the solid through a small funnel. Stopper the burette

and invert it several times to mix the contents. Continue adding the solid until it has all dissolved.

WEAR EYE PROTECTION

- 9 Record the volume of the solution.
- 10 Repeat the procedure in steps 7–9 for the other two solids.
- **11** Calculate the volume of the solid you added in each case, using its mass and the density given below.

Solid	Density/g cm ⁻³
sodium chloride	2.2
calcium chloride	2.5
iron(III) chloride	2.8

12 Use the relationship

volume change = (volume of solution) – (volume of solid + volume of water)

to work out in each case the change in volume when the solid dissolves.

13 Draw up a table to show your results.

Interpreting your results

You may wish to revise ideas about the hydration of ions in **Chemical Ideas 3.2** and **5.1** to help you with this part of the activity.

- **a** Interpret the enthalpy changes measured in Part 1 in terms of the processes that occur when the solids dissolve and the interactions between the chemical particles present.
- **b** Use these ideas to explain the volume changes you noted in Part 2 of the activity.



Ol.3 What factors affect the enthalpy change of formation of an ionic compound?

In this activity you will see how the formation of an ionic compound from its constituent elements can be imagined to take place in several stages, each with its own enthalpy change. The values of these enthalpy changes determine the overall value for the enthalpy change of formation of the compound.

Part 1: Constructing an entbalpy cycle

The standard enthalpy change of formation of sodium chloride, $\Delta H_{f}^{\Theta}(\text{NaCl}(s))$, is the enthalpy change for the reaction:

 $Na(s) + \frac{1}{2}Cl_2(g) \rightarrow NaCl(s)$ $\Delta H_f^{\oplus}(NaCl(s)) = \text{ standard enthalpy change}$

of formation of sodium chloride

You can imagine the stages involved as:

- (1) turning the elements Na(s) and $Cl_2(g)$ into gaseous atoms
- (2) turning the atoms into gaseous ions
- (3) bringing the gaseous ions together to form NaCl(s)

The following equations describe the stages:

- (a) $Na(s) \rightarrow Na(g)$
- (b) $Na(g) \rightarrow Na^+(g) + e^-$
- (c) $\frac{1}{2}Cl_2(g) \rightarrow Cl(g)$
- (d) $Cl(g) + e^{-} \rightarrow Cl^{-}(g)$
- (e) $\operatorname{Na}^+(g) + \operatorname{Cl}^-(g) \rightarrow \operatorname{NaCl}(s)$
- **1** Draw up a table and match the equations (a) to (e) with the stages (1) to (3) above.
- **2** Each of the equations (a) to (e) defines a standard enthalpy change. Match the enthalpy changes (i) to (v) below with the equations (a) to (e).

(i)	lattice enthalpy of sodium chloride	$\Delta H_{LE}^{\bullet}(\text{NaCl}(s))$
(ii)	first electron affinity of chlorine	$\Delta H_{\rm EA}^{\bullet}$ (Cl(g))
(iii)	standard enthalpy change of atomisation of sodium	$\Delta H_{at}^{\bullet}(Na(s))$
(iv)	standard enthalpy change of atomisation of chlorine	$\Delta H_{at}^{\Theta}(\frac{1}{2}Cl_2(g))$
(v)	first ionisation enthalpy of sodium	$\Delta H_{i}^{\Theta}(1)$ (Na(g))

- **3** Draw out the enthalpy cycle showing how you get from the elements in their standard states, Na(s) and $\frac{1}{2}Cl_2(g)$, to the compound, NaCl(s), using the stages defined by the equations (a) to (e). Label each stage with the appropriate symbol taken from (i) to (v). This type of enthalpy cycle for a simple ionic compound is known as a **Born-Haber** cycle.
- **4** Apply Hess's Law to the cycle and obtain an equation that shows how $\Delta H_{\rm f}^{\,\Theta}({\rm NaCl}({\rm s}))$ depends on the other enthalpy changes.
- **5** Set up a spreadsheet that will calculate ΔH_{f}^{e} for any Group 1 halide. The only data input required will be the values for the enthalpy changes (i) to (v).

Part 2: Using the cycle

6 Use bar charts, from your spreadsheet, to explore how ΔH_{f}° depends on ΔH_{LE}° and $\Delta H_{i}^{\circ}(1)$ for the series of compounds:

i NaF, NaCl, NaBr, NaI

ii LiCl, NaCl, KCl, RbCl

Comment on any patterns you see.

A2 LEVEL

QUESTION

How would you need to adapt your enthalpy cycle and spreadsheet to cope with the halides of Group 2 elements?



The first part of this activity is optional. It introduces you to some practical techniques which you have not previously encountered in the course, and allows you to measure a value for the enthalpy change of vaporisation (ΔH_{vap}) of water. You can then compare this with values for some other liquids and use the ideas you have learned in Chemical Ideas 4.4 to interpret the results.

Part 1: Measuring the enthalpy change of vaporisation of water (Optional extension)

Requirements

- 12Vd.c. supply
- joulemeter
- electrical heating coil (see Figure 1)
- apparatus for boiling and condensing water (see Figure 1)
- small beakers or specimen tubes (2)
- access to balance

CARE Eye protection (must be worn.



What you do_

- **1** Assemble the apparatus shown in Figure 1, for boiling and condensing water. Use water close to boiling in the side-arm boiling tube.
- **2** Position one small beaker as shown in the diagram; have a second weighed beaker ready.



- **3** Switch on the power. Allow the water to warm up and then boil for a few minutes until it drips at a slow steady rate from the condenser into the unweighed small beaker. Then change over the beakers and simultaneously reset the joulemeter.
- 4 Allow about 3 cm^3 of water to distil over, then exchange the beakers once more, reading the joulemeter at the same instant.
- **5** Record the mass of beaker plus water and calculate the mass of water collected.



Figure 1 Apparatus for measuring the enthalpy change of vaporisation of water

Part 2: Looking at your results

You should do this part of the activity even if you have not done the practical in Part 1. The questions which follow should encourage you to apply what you have learned in **Chemical Ideas 4.4**. Make sure you have read these ideas before you attempt the questions.

- **a** Calculate a value for ΔH_{vap} for water in kJ mol⁻¹. (In a typical student's experiment, 8100 J were needed to distil 2.95 g of water.)
- **b** Explain why the boiling tube is surrounded with expanded polystyrene but the condenser is left uncovered.
- **c** Before any water was collected, the water was allowed to boil for several minutes until drops fell steadily from the condenser. Explain why this helps you achieve greater accuracy in your estimate of ΔH_{vap} .

Table 1 lists values of ΔH_{vap} and boiling point (T_{b}) for water and four organic compounds.

Liquid	Formula	$\Delta H_{ m vap}/ m kJ~mol^{-1}$	<i>Т</i> _b /К	
trichloromethane	CHCI ₃	+29.3	335	
hexane	C_6H_{14}	+28.8	342	
cyclohexane	$C_{6}H_{12}$	+30.1	354	
methylbenzene	C ₇ H ₈	+33.4	384	
water	H ₂ O	+40.6	373	

Table 1 Entbalpy changes of vaporisation (ΔH_{vap}) and boiling points (T_b) for some compounds

- **d i** Water has a high ΔH_{vap} value compared with the other substances in Table 1. Use your knowledge of intermolecular forces to explain this observation.
 - ii Ethanol also has an unusually high ΔH_{vap} value of $+38.5 \text{ kJ mol}^{-1}$. Explain how the changes which occur during boiling are similar for ethanol and water.
- **e** All the ΔH_{vap} values are highly endothermic. Nevertheless, the compounds are volatile and can easily be boiled. What other change which favours the process accompanies vaporisation?
- **f i** At the boiling point, a liquid and its vapour are in equilibrium. What can you say about the total entropy change for a system which is at equilibrium?
 - ii What is the relationship between the entropy change on vaporisation of a liquid (ΔS_{vap}) and the entropy change in the surroundings (ΔS_{surr}) at the boiling point?
 - iii ΔS_{surr} during an endothermic change can be calculated from the thermal energy supplied and the temperature at which the change occurs. For boiling, the relationship is:

$$\Delta S_{\rm surr} = \frac{-\Delta H_{\rm vap}}{T_{\rm b}}$$

Use this relationship to calculate ΔS_{surr} values for the five compounds in Table 1.

- **g i** Use your answers for **f ii** and **f iii** to calculate ΔS_{vap} values for the five compounds in Table 1.
 - ii What general observation can you make about the values of ΔS_{vap} for the four organic compounds in Table 1?
 - iii Explain this observation in terms of the molecular-kinetic model which chemists use to describe the structures of liquids and vapours.
 - iv How does the value of ΔS_{vap} for water compare with the value you found for the four organic liquids? Explain why there is a difference.
- **h i** Convert the values for ΔH_{vap} in Table 1 into kJ kg⁻¹.
 - ii These figures should show you that, for the transport of a given mass of vapour through the atmosphere, water carries much more energy with it than the other substances. Explain why this is important for the climate of the Earth.



What crystals form when a solution is cooled?

Freezing and crystallisation cause density changes in the oceans and are among the factors that drive ocean currents. This activity uses copper(II) sulphate, which is easy to distinguish from ice crystals, to investigate the effect of cooling solutions containing different concentrations of an ionic compound. The questions, which help you to interpret your observations, also provide an opportunity for you to reinforce what you have learned in Chemical Ideas 4.4 about entropy changes and their relationship to enthalpy changes and temperature.

CARE Eye protection

must be worn.

Requirements_

- saturated copper(II) sulphate solution (4 cm³)
- 250 cm³ beaker
- test-tubes (2) and rack
- crushed ice
- table salt
- -10°C-110°C thermometer

What you do.

- **1** Make an ice–salt freezing bath by mixing crushed ice and table salt in a beaker.
- **2** Place about 3 cm³ of saturated copper(II) sulphate solution in one test-tube. Add 1 cm³ of saturated copper(II) sulphate solution to 2 cm³ of water in the other test-tube.
- **3** Place the test-tubes into the freezing bath and allow the solutions to cool. Make a note of the temperature at which crystals form in each solution, the appearance of the crystals, and whether they float or sink.
- **4** Before you throw away the contents of the freezing bath, make a note of its temperature and appearance.

What it means

- a When a solution of copper(II) sulphate cools, what determines whether it is water or copper(II) sulphate that crystallises out first?
- **b** Describe how you think the density of a copper(II) sulphate solution changes when:
 - i copper(II) sulphate crystals form
 - ii ice crystals form.
- **c** The salt concentration in sea water is typically about 0.5 mol dm⁻³. How might ocean currents be affected if the sea were saturated with salt?
- **d** The enthalpy change when water freezes to form ice is –6010 J mol⁻¹. What will be the gain in entropy of the surroundings when water freezes at 273 K?
- At 273 K the entropy of ice is lower than the entropy of water by 22.0 J K⁻¹ mol⁻¹. Explain why the entropy of ice is lower.
- f Comment on the total entropy change of the system and surroundings that accompanies the freezing of water at 273 K.

- **g** The enthalpy change (ΔH) when pure water freezes to produce ice is the same as the enthalpy change when copper(II) sulphate solution freezes to produce ice. However, ice crystals do not form from copper(II) sulphate solution until the temperature has fallen below 273 K. This is because the entropy changes (ΔS_{sys}) are different for the two processes.
 - i For which process will the entropy change $(\Delta S_{_{Sys}})$ be greater:
 - pure water \rightarrow ice at its freezing point; or copper(II) sulphate solution \rightarrow ice at its freezing point?

Explain your answer.

ii At the freezing point, ice is in equilibrium with the water or solution which is freezing. Therefore ΔS_{total} must be zero. You have made decisions about ΔS_{sys} in **g i**. How must ΔS_{surr} compare for the two processes:

pure water \rightarrow ice at its freezing point; and copper(II) sulphate solution \rightarrow ice at its freezing point?

iii The entropy change in the surroundings (ΔS_{surr}) is related to ΔH and the temperature at which freezing occurs. Use this relationship to explain why ice forms from copper(II) sulphate solution at a temperature below 273 K.

The same idea explains why it is possible to have water, ice and salt present together at temperatures well below 273 K. So you can melt ice on roads in winter, and make it run off, by adding salt. You can also make freezing baths like the one you used in this activity.



WEAR EYE PROTECTION

Finding out more about weak acids

04

You can compare some of the properties of weak and strong acids by using ethanoic acid as a typical weak acid and hydrochloric acid as a typical strong acid. The measurements you make can be interpreted in terms of the theory in Chemical Ideas 8.2.

By the end of the activity you should have seen that pH is less sensitive to changes in weak acid concentration than strong acid concentration. You should then appreciate one of the reasons why the thousandfold change in atmospheric carbon dioxide concentration that has occurred during the Earth's history has not caused a major change in the pH of the oceans.

Requirements

- pH meter
- 100 cm³ beaker
- solutions of hydrochloric acid of the following concentrations (about 50 cm³ of each solution):
 - $-0.1 \,\mathrm{mol}\,\mathrm{dm}^{-3}$
 - $-0.03 \,\mathrm{mol}\,\mathrm{dm}^{-3}$
 - $-0.01 \,\mathrm{mol}\,\mathrm{dm}^{-3}$
 - $-0.003 \, mol \, dm^{-3}$
 - $-0.001 \, mol \, dm^{-3}$
- solutions of ethanoic acid of the same concentrations as above (about 50 cm³ of each solution)

What you do_

- **1** Pour sufficient 0.001 mol dm⁻³ hydrochloric acid into a beaker to allow you to measure the pH. Record the pH and return the solution to the stock bottle.
- **2** Repeat the measurements with the other hydrochloric acid solutions. If you work from the lowest to the highest concentration you can use the same beaker, and there is no need to rinse it out or clean the glass electrode between readings.
- **3** Now repeat steps **1** and **2** using the ethanoic acid solutions. Before you start, make sure the glass electrode is rinsed with distilled water to remove all traces of hydrochloric acid.
- 4 Present all your results in the form of a table.

Discussion of results

A2 LEVEL

- a i Explain what you understand by the term pH.
 ii Explain how changes in pH and changes in [H⁺(aq)] are related.
- **b** From your results, what appears to be the pH change associated with:
 - i a 10-fold change in the concentration of hydrochloric acid?
 - ii a 100-fold change in the concentration of ethanoic acid?
- **c i** Explain what you understand by the term *strong acid*. How is this different from a *concentrated acid*?
 - ii Explain why you would expect a 10-fold change in the concentration of a strong acid to lead to a 10-fold change in [H⁺(aq)].



- iii How does your answer to **b** i compare with the behaviour you would expect of a strong acid?
- **d i** Explain what you understand by the term *weak acid*.
 - **ii** Write an expression for the acidity constant (K_a) of ethanoic acid. (You can abbreviate the formula to HA for convenience.)
 - iii It is reasonably accurate to regard [H⁺(aq)] and [A⁻(aq)] as being equal in the solutions of ethanoic acid used in this activity. Explain why we can do this.
 - iv Another reasonably accurate assumption is to regard [HA(aq)] as equal to the amount of acid used to make 1 dm³ of each solution. For example, we can say [HA] = 0.1 mol dm^{-3} in ethanoic acid solution of concentration 0.1 mol dm^{-3} . Explain why this is an accurate assumption.
 - Rewrite K_a for ethanoic acid solutions of concentrations 0.1 mol dm⁻³ and 0.001 mol dm⁻³, using the assumptions discussed in d iii and iv. You should be left with expressions for K_a which involve only [H⁺(aq)] and numbers.
 - **vi** How does the value for [H⁺(aq)] change when a 0.1 mol dm⁻³ solution of ethanoic acid is diluted by a factor of 100, to 0.001 mol dm⁻³?
 - vii You have just worked out how you would expect [H⁺(aq)] and [HA(aq)] to be related for ethanoic acid – an example of a weak acid. How does the observation you made in **b** ii compare with this expected behaviour?

04.2

Investigating some buffer solutions In this activity you make pH measurements of some buffer solutions and see what bappens when you add water, acid and alkali to them. You then compare their behaviour to a solution of similar pH that is not a buffer. Interpreting the results should reinforce the ideas you learned about in Chemical Ideas 8.3.

Requirements.

- $0.5 \,\mathrm{mol}\,\mathrm{dm}^{-3}$ solutions of the following:
 - ethanoic acid (75 cm^3)
 - potassium (or sodium) ethanoate (75 cm³)
 - methanoic acid (25 cm^3)
 - potassium (or sodium) methanoate (25 cm³)
 - ammonium chloride (25 cm^3)
 - ammonia solution (25 cm^3)
 - hydrochloric acid (15 cm^3)
 - potassium (or sodium) hydroxide (15 cm³)
- 25 cm³ measuring cylinder
- 10 cm³ measuring cylinder
- distilled-water wash bottle
- 100 cm³ beakers (7)
- pH meter
- 1×10^{-4} mol dm⁻³ nitric(V) acid (50 cm³)
- glass rod

What you do.

As you work through this activity, fill in your results in Table 1.

- **1** Make up three buffer solutions by mixing 25 cm³ portions of the 0.5 mol dm⁻³ solutions in the pairs listed below. (Use 100 cm³ beakers to hold the buffer solutions and stir each one with a glass rod.)
 - **Buffer A** ethanoic acid + potassium ethanoate
 - Buffer B methanoic acid + potassium methanoate
 - Buffer C ammonium chloride + ammonia solution

Measure the pH of each buffer using a pH meter. Rinse the glass electrode with distilled water between measurements.

- **2** Remove 5 cm³ of Buffer A and add it to 45 cm³ of distilled water in a fourth beaker. This is **Buffer D**. Record its pH.
 - a How are the pH values of Buffer A and Buffer D related?
 - **b** Does the pH of a buffer depend on the *total* amounts of acid and salt present or on the *ratio* of their amounts? Explain your answer.
- **3** Make up two more buffers using 0.5 mol dm⁻³ ethanoic acid and 0.5 mol dm⁻³ potassium ethanoate as follows:

Buffer E 10 cm^3 ethanoic acid + 40 cm^3 potassium ethanoate**Buffer F** 40 cm^3 ethanoic acid + 10 cm^3 potassium ethanoate

Record the pH values.

c Use Le Chatelier's principle to explain the way the pH values vary among the Buffers A, E and F.



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The measurements you have made show that buffers can have pH values spread over a wide range. There are two steps to designing a buffer with a particular pH:

- 'coarse tuning' to select the pH region
- 'fine tuning' to adjust the pH to the actual value required.

You have seen that the pH of a buffer depends on:

- the chemical system chosen
- the proportions of acid and salt used.
- d Which of these two factors would you make use of in
 - i the 'coarse tuning'
 - ii the 'fine tuning'?
- **4** Place 50 cm^3 of $1 \times 10^{-4} \text{ mol dm}^{-3}$ nitric(V) acid into another beaker. This is **Solution G**; it has a similar pH to the ethanoic acid + potassium ethanoate mixtures, but it is not a buffer solution.

Add 3 cm^3 of 0.5 mol dm^{-3} hydrochloric acid to solutions A, D and G. Record the pH changes.

- **e** Explain the different effects of adding the same amount of hydrochloric acid to these three solutions.
- **5** Add 3 cm³ of 0.5 mol dm⁻³ hydrochloric acid or potassium hydroxide to buffers B, C, E and F. You are free to choose which solutions to add to which buffers. Record the pH changes.
 - f Explain the effects of adding acid or alkali to the buffers.

Solution	System	рН	∆pH + 3 cm ³ HCI	∆pH + 3 cm ³ KOH	Was the solution an effective buffer?
A					
В					
С					
D					
E					
F					
G					

Table 1 Results table

A2 LEVEL



This activity belps you get your notes in order at the end of this unit.

Use this list as the basis of a summary of the unit by collecting together the related points and arranging them in groups. Check that your notes cover the points and are organised in appropriate ways.

Most of the points are covered in the **Chemical Ideas**, with supporting information in the **Storyline** or **Activities**. However, if the *main* source of information is the Storyline or an Activity, this is indicated.

- Factors that determine the relative solubility of a solute in aqueous and non-aqueous solvents.
- The meaning and use of the terms: *enthalpy change of solution*, *lattice enthalpy*, *enthalpy of solvation* (*bydration*).
- The solution of an ionic solid in terms of an enthalpy cycle involving enthalpy change of solution, lattice enthalpy and enthalpies of solvation (hydration) of ions.
- The trends in solubility in water of the hydroxides and carbonates of the elements of Group 2 (excluding Be) in terms of lattice enthalpies and enthalpies of hydration of the ions.
- The Born-Haber cycle for simple ionic compounds.
- Entropy as a measure of the number of ways that molecules and their associated energy quanta can be arranged.
- The process of dissolving in terms of energy and entropy factors.
- The tendency of a process to occur in terms of entropy changes in the system (ΔS_{sys}) and surroundings (ΔS_{surr}), and the requirement that the total entropy change (ΔS_{total}) should be positive.

- Calculations of entropy changes using the expression: $\Delta S_{\text{total}} = \Delta S_{\text{sys}} + \Delta S_{\text{surr}}.$
- Comparison of the following properties of water with those of other liquids, and other hydrides of Group 6 elements, and the relationship of these properties to molecular structure: specific heating capacity, enthalpy change of vaporisation, and density changes on melting.
- The influence of oceans on climate in terms of the characteristic properties of water (**Storyline O3**).
- The meaning and use of the following terms: *strong acid*, *strong base*, *pH*.
- The ionic product of water, K_{w} .
- Calculation of the pH of solutions of strong acids and strong bases.
- The meaning and use of the following terms: *weak acid*, *acidity constant K*_a, pK_a.
- Calculation of the pH of solutions of weak acids.
- · How buffer solutions work, and their applications.
- Calculation of the pH of a buffer solution.
- The meaning and use of the term *solubility product* for simple ionic compounds of formula X^{*n*+}Y^{*n*-}.
- The use of solubility products to perform calculations concerning dissolving and precipitation processes.
- Acid–base and precipitation processes in the oceans in terms of K_a and K_{sp} (Storyline O4).
- The global influence of the processes occurring when carbon dioxide dissolves in water (**Storyline O4**).