Teaching ideas for Chapter 7, *Equilibrium*

Questions

Two worksheets of questions are provided:
- the first worksheet deals with the Standard Level part of the syllabus
- the second worksheet is for Higher Level only.

There are also a large number of questions available in the Coursebook and on the accompanying CD-ROM.

Teaching ideas

- The idea of dynamic equilibrium could be discussed with regard to a shop in which the shelves are filled at the same rate as they are emptied or a bath/lake where the water is flowing in/out at the same rate.
- The idea of dynamic equilibrium can be illustrated in various ways.
  - One student could be given 5 balls and another 3 balls. If the students continuously pass 1 ball to each other then the number of balls each has remains constant.
  - Alternatively, two troughs can be set up – one containing water and the other empty. Two students are given a tin can (or other suitable container) each, of equal sizes, with a hole bored in the side near the bottom. The first student puts the can into trough 1 (the one containing water), allows the can to fill with water, covers the hole with their finger and transfers the water to trough 2 (which is the empty trough). The second student puts their can into trough 2, allows water to enter the can to the same depth as in the trough, puts their finger over the hole and transfers the water to trough 1. This is repeated until it is clear that the water levels in the two troughs have become and are remaining fairly constant. The key points are that the cans are of equal sizes and water is only transferred equivalent to the depth of water in any one trough. Various starting amounts of water in each trough could be investigated.
    http://www.saskschools.ca/curr_content/chem30/modules/module5/lesson1/dequilibrium.htm
    http://www.svusd.org/hp_images/3313/D18544-aquariumequilibrium.pdf
- Students can research the conditions used in various industrial processes and relate these to what they have learnt about rates of reactions and Le Chatelier's principle. The factors that determine the conditions used in industrial processes could be discussed. To what extent do environmental considerations influence those conditions?
- Students could be encouraged to read Le Chatelier's original paper: http://web.lemoyne.edu/~giunta/paperabc.html
- Ammonia and sulfuric acid are both used to make fertilisers. The importance of artificial fertilisers could be discussed. To what extent are artificial fertilisers essential in certain parts of the world for preventing famine? Why has there been a rise in organic farming over the last few years? Are there any health and environmental effects associated with artificial fertilisers?
- The use of ammonia to make explosives could also be discussed. Students could research how the development of the Haber–Bosch process for the production of ammonia prolonged World War I.
- Students could research the life of Fritz Haber or a debate could be held debating whether he was a hero or a villain.

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- The uses of sulfuric acid could be discussed as well as how much is produced in different countries. Is there a link between the scale of sulfuric acid production and the state of development of a country?
- A use of sulfuric acid that students are familiar with is in car batteries. They could look at the environmental effects of disposing of car batteries.
- Equilibrium constants could be further related to ΔG for more advanced students. $\Delta G = 0$ at equilibrium. A change in conditions will cause ΔG to no longer be zero. The reaction will have ΔG negative in one direction and therefore will move in that direction to a new position of equilibrium.

Practical activities

Safety

Extreme care must be exercised when carrying out any practical activities in the classroom and a risk assessment should be conducted before carrying out the experiments.

Demonstrations

There are various demonstrations that can be carried out to demonstrate how changing conditions affect the position of equilibrium.

- The effect of pressure can be investigated using NO_2/N_2O_4 in a sealed gas syringe.
- The effect of temperature can be investigated using a mixture of $CoCl_2(aq)$ and concentrated HCl (**Care!**). Putting the mixture into ice water gives a pink colour and heating it in a Bunsen flame gives a blue colour.
- The effect of temperature can also be investigated by using NO_2/N_2O_4 sealed in a boiling tube.
- The effect of concentration can be investigated by adding acid and alkali to an acid/base indicator or, for a more spectacular demonstration, by adding concentrated hydrochloric acid and concentrated ammonia (**Care!**).
- The effect of concentration can also be investigated by adding acid and alkali to potassium dichromate(VI) solution.
- Further details on these demonstrations are given below:
  http://www.saskschools.ca/curr_content/chem30_05/3_equilibrium/teacher/demos_equilibrium.pdf
  http://www.practicalchemistry.org/experiments/31st-jan-the-equilibrium-between-coh2o62-and-cocl42-in-aqueous-solution,178,EX.html
  http://www.union.edu/academic_depts/chemistry/faculty/fox/Chemical%20Demonstrations/docs/11%20Cobalt%20Complex%20Equilibrium.pdf
  http://media.rsc.org/Classic%20Chem%20Demos/CCD-81.pdf
  http://chemed.chem.purdue.edu/demos/demosheets/16.3.html
  http://chemed.chem.purdue.edu/demos/main_pages/16.2.html
  http://www.practicalchemistry.org/experiments/advanced/equilibrium/topic-index.html
- Phase equilibrium may be illustrated using bromine in a gas jar in a fume hood (**Care!**). A small amount of liquid bromine is introduced to the gas jar and the lid put on. This is left for some time until the depth of colour of the vapour remains constant.

Student practicals

- A classic experiment on this part of the syllabus is to determine an equilibrium constant for a reaction.

  http://www.chemtopics.com/aplab/eqetac.pdf
  http://www.files.chem.vt.edu/RVGS/ACT/lab/Experiments/Exp_9-Equil.html

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- Other practicals can be found at:

  http://www.calpoly.edu/~cbailey/125LabExperiments/Equilibrium/EquilExpIntro.html
  http://taylord.people.cofc.edu/Experiment%205A.pdf
  http://www.uccs.edu/~chemistry/nsf/106%20Expt3V-Equilibrium.pdf
  http://web.centre.edu/shiba/che132L/equil_const.pdf

Common problems

- Students often have difficulties accepting that K_c should be unaffected by changes in pressure or concentration. One way of convincing students how this can happen is to use quantitative examples, showing how a change in the equilibrium concentrations can still result in K_c having the same value. The example below could be used with students.

The effect of pressure

If there are different numbers of gaseous molecules on either side of the equation then a change in pressure will change the position of equilibrium but not the value of the equilibrium constant.

Consider the following data for the $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$ reaction at 500 K:

$[N_2]$	$[H_2]$	$[NH_3]$	$\%NH_3$	Total P / atm	K_c / mol^{-2} dm^6
0.72	0.72	0.41	22	7.7	0.62
0.72	0.96	0.63		9.6	
0.48	1.44	0.94		11.9	
0.72	2.17	2.14		20.9	
1.20	12.0	35.9		204	
4.81	12.0	71.7	81	368	

All concentrations are measured at equilibrium.

As can be seen from this data, the value of K_c remains constant to within experimental error even though the position of equilibrium changes markedly as the pressure is increased.

This data also illustrates a danger of relying too much on K_c to predict extent of reaction. The value of K_c remains constant but the yield of ammonia changes from 22% to 81%.

How can the effect of increasing the pressure be related to Le Chatelier's principle?

If the first of the equilibrium mixtures above is compressed so that all concentrations are doubled (same number of moles in half the volume), the value of the ratio

$$\frac{[NH_3]^2}{[N_2][H_2]^3} \text{ becomes } \frac{0.82^2}{1.44 \times 1.44^3} = 0.16 \text{ mol}^{-2} \text{ dm}^{-3}$$

but, at this temperature, the value of K_c is 0.62 mol^{-2} dm^{-3}. Therefore, the system is no longer at equilibrium as the ratio is too low. Nitrogen and hydrogen must react together to give more ammonia and so increase the value of the ratio.

In this particular case the new concentrations at equilibrium become:

$[N_2]$ = 1.33 mol dm^{-3}
$[H_2]$ = 1.10 mol dm^{-3}
$[NH_3]$ = 1.04 mol dm^{-3}

and the percentage of ammonia has gone up to 30%, which is in accordance with Le Chatelier's principle:

When a system at equilibrium is subjected to a change in pressure or concentration the position of equilibrium will be shifted in order to restore the value of the equilibrium constant.

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ICT

There are many opportunities for using IT in this topic. Many of the excellent websites are listed below contain simulations.

- Equilibrium and Haber process simulations:
  http://www.freezeray.com/chemistry.htm
- Equilibrium animations:
  http://www.chem.iastate.edu/group/Greenbowe/sections/projectfolder/animationsindex.htm
- Equilibrium simulations
  http://www.chm.davidson.edu/ronutt/che115/equkin/equkin.htm
  http://chemconnections.org/Java/equilibrium/index.html
  http://ccl.northwestern.edu/netlogo/models/ChemicalEquilibrium
  http://phet.colorado.edu/en/simulation/soluble-salts
  http://web.mst.edu/~gbert/EQUIL/equil.html
- Haber process spreadsheet:
  http://chemistry.compendiarious.net/Haber's%20ProcessCalc.xls
- Vapour pressure:
  http://www.chem.iastate.edu/group/Greenbowe/sections/projectfolder/flashfiles/propOfSoln/vp3.html
  http://www.mhhe.com/physsci/chemistry/essentialchemistry/flash/vaporv3.swf
- Equilibrium spreadsheet:
  http://www.molsci.ucla.edu/source_mat/KeqVsExtentOfReaction,v12.xls
- Students could also set up a spreadsheet to calculate equilibrium concentrations and equilibrium constants.

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