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Investigating High School Students' Understanding of Chemical Equilibrium Concepts

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This study investigated the year 12 students' (N = 56) understanding of chemical equilibrium concepts after instruction using two conceptual tests, the *Chemical Equilibrium Conceptual Test 1(CECT-1)* consisting of nine two-tier multiple-choice items and the *Chemical Equilibrium Conceptual Test 2(CECT-2)* consisting of four structured questions. Both these tests were administered before and after the intervention. Students' responses to the items in both the instruments indicated limited understanding of the various concepts related to chemical equilibrium. Less than 50% of the students provided correct responses to four of the nine items in the *CECT-1.* The total scores in the *CECT-1* ranged from 0 to 8 with a mean score of 4.14 (out of a maximum of 9). In the *CECT-2* the total scores ranged from 7 to 17 with a mean score of 11.0 out of a maximum score of 22. Almost half the number of students (44.6%) scored less than 50% of the total marks in the *CECT-2*; only 0% to 42.9% of students scored the maximum possible marks for each of the four items while achievement in all four items of the *CECT-2* was below 50%. The findings will be valuable and assist teachers in planning their instruction on chemical equilibrium by taking into consideration students' preconceptions about the topic.

Keywords: chemical equilibrium, dynamic equilibrium, Le Chatelier's Principle, reversible reactions

INTRODUCTION

The topic of chemical equilibrium has been widely researched at secondary school and university levels in several countries since the 1960s, for example, in

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Australia (Tyson, Bucat & Treagust, 1999), the US (Piquette & Heikkinen, 2005), Spain (Quilez-Pardo & Solaz-Portoles, 1995), The Netherlands (van Driel et al., 1998), China (Cheung, 2009), India (Banerjee, 1991) and Turkey (Özmen, 2008). The concept is important in chemistry as it is essential for the understanding of several other concepts like oxidation and reduction, phase changes, solubility and weak and strong acids (Voska & Heikinnen, 2000). This research study evaluated the understanding of chemical equilibrium concepts held by students from a private secondary school in the first year of a two-year pre-university course in Malaysia where science was taught in English for the last time. (From 2014, instruction in science will revert to the Malay language, the national language of the country after having trialled instruction in science using English for the past several years). Hence, this study is a last opportunity to collect data and analyze learning by Malaysian students in the medium of English instruction and may serve as a useful baseline for future studies of chemistry when Malaysian students are taught in the medium of the Malay language. Furthermore, this change in instructional medium is not unique to Malaysia. For example, as a result of recent school reforms in Qatar's public schools and largest university, these institutions have reverted from English to Arabic as the medium of instruction (Khatri, 2013).

The items in the conceptual instruments used in this study that were adapted from previous studies or developed by the authors will enhance the research literature with conceptual items that could be referred to by other researchers as the topic of chemical equilibrium is widely taught in the curricula of several developed and developing countries around the world.

Theoretical background

Chemical equilibrium is a difficult concept to understand for both teachers and students partly because of its abstract nature (Quiliz-Pardo & Solaz-Portoles, 1995). Students have been found to provide correct answers to questions on chemical equilibrium but were unable to provide correct reasons (Quiliz-Pardo & Solaz-Portoles, 1995).

Since the 1960s, chemical equilibrium, considered one of the most difficult topics in secondary school chemistry (Finley et al., 1982; Bergquist & Heikkinen, 1990), has been extensively researched because the topic is included in several high school and tertiary chemistry curricula (Tyson, Bucat & Treagust, 1999; Voska & Heikkinen, 2000; Kousathana & Tsaparlis, 2002; Quilez, 2004; Özmen, 2008; Cheung, 2009; Cheung et al., 2009). Research has shown that high school chemistry students' conceptions of equilibrium reactions were influenced by their prior experience of reactions that proceeded to completion (Hackling & Garnett, 1985; Pedrosa & Dias, 2000), resulting in students displaying alternative conceptions such as (1) the rate of the forward reaction increases from the commencement of the reaction until equilibrium is reached, (2) there is a simple arithmetic relationship between the concentrations of the reactants and products, and (3) when a change is made to a system at equilibrium (e.g. addition of a reactant), the rate of only the forward reaction increases, while that of the reverse reaction decreases. It is important to know about students' alternative conceptions about chemical equilibrium so that appropriate remedial measures may be implemented to equip them to better cope with related topics at more advanced levels (Thomas & Schwenz, 1998).

To address these learning difficulties, Grade 10 students (15-16 year-olds) in a study were introduced to the basic concepts of chemical equilibrium by challenging students' previously acquired conceptions of chemical reactions (Van Driel et al., 1998). The researchers found the need to change several conceptions held by students about chemical equilibrium. These included addressing the concept of the reversibility of chemical reactions in order that students were able to accept the "incompleteness of a chemical conversion as an empirical fact" (p. 389). In doing so, they suggested introducing the idea of *dynamic equilibrium* to account for the reversibility of chemical reactions and for the reactions not going to completion. So, conceptual change among the students through cognitive dissonance in an experimental course was effected by suggesting the possibility of reversibility of chemical reactions, the possibility of reactions not going to completion and the two opposite chemical reactions that are taking place existing in dynamic equilibrium. The researchers were cognisant of the conditions necessary for conceptual change requiring students to first become dissatisfied with their existing conceptions about chemical reactions, and then accepting the new conception if it seemed intelligible, plausible and fruitful (Posner et al., 1982). In addition, influences like social and affective factors (Pintrich et al., 1993) were considered using experiments (e.g., the equilibrium mixture containing deep blue cobalt(II) tetrachloro and pink cobalt(II) hexahydrate complexes) in which colourful changes that occurred kept students interested and encouraged them to interact with each other to arrive at acceptable explanations. The study found that the experiments caused students to be dissatisfied with their existing conceptions of chemical reactions and they generally accepted the concept of dynamic equilibrium as a model to account for the conflicting conceptions that they originally held about chemical equilibrium reactions.

A major issue in the study of chemical equilibrium is the use of Le Chatelier's Principle (LCP) and its application in textbooks (Canagaratna, 2003). Often expressed as "If a chemical system is subjected to a perturbation, the equilibrium will be shifted such as to partially undo this perturbation" (Torres, 2007; p. 516), this principle, because of its ambiguity, often leads to erroneous predictions. Ignorance about the limitations of LCP among teachers often has led to teachers' and students' inappropriate use of the principle (Quilez-Pardo & Solaz-Portoles, 1995).

In expressing the difficulties associated with the teaching and learning of the chemical equilibrium topic, Tyson, Bucat and Treagust (1999) have recognised the need to understand a multifaceted approach involving (1) Le Chatelier's Principle, (2) the equilibrium law, and (3) an analysis of the forward and reverse reaction rates using the collision theory, that students could use when explaining the effects of disturbing a system in chemical equilibrium. In extended interviews with three Year 12 students, the researchers found that not all students used LCP or the equilibrium law in making successful predictions about the equilibrium changes; their choice of explanation depended on the context of the problem.

To gain insight into a specific case, we studied the teaching and learning of the chemical equilibrium topic based on the Malaysian Higher School Certificate (HSC) syllabus for 2012/2013 (Malaysian Examinations Council, 2011).

Objectives and research question

The main objective of this research was to elucidate Year 12 students' understanding of chemical equilibrium concepts. Using this information, teachers would be able to modify their classroom instruction to better facilitate understanding of chemical equilibrium concepts by their students. This study is also relevant to the teaching and learning of chemical equilibrium concepts in general as the topic is included in the science curricula of several other countries. The findings of this study will therefore, benefit science teachers and other educators in countries where the topic is taught. The study was guided by the main research question: To what extent do the two conceptual instruments probe students' understanding of chemical equilibrium concepts?

METHODOLOGY

Research design

The research used a quantitative design (Cohen et al., 2011) to ascertain Year 12 students' understanding of chemical equilibrium concepts after about eight hours of instruction over three weeks. As the study involved elucidating students' conceptual understanding of chemical equilibrium concepts and not the efficacy of the instructional program, we did not see the necessity of involving a comparison group in the study.

Research sample

The students who were involved in the study constituted a convenience sample of 56 high-achieving Lower 6 (year 12) students from two classes in a private secondary school who were taught by the same teacher who had agreed to the research being conducted in her classes. Involving the same chemistry teacher ensured consistency in instruction in the two classes. Prior consent was obtained by the first author from the students who had agreed to participate in the study.

Instructional program

1. Beginning with the reactants, as a reaction mixture approaches equilibrium, the concentrations of the reactants decrease while that of the products increase.
2. Beginning with the reactants, as a reaction mixture approaches equilibrium, the rate of the forward reaction decreases and the rate of the reverse reaction increases.
3. When equilibrium has been established, the concentrations of all the reactants and products remain constant.
4. Once equilibrium has been established, the forward and reverse reactions continue to occur (i.e. the equilibrium system is a dynamic one, not static), at the same rate.
5. At equilibrium, the concentrations of the reactants and products are related by the equilibrium law, where K_c and K_p are the equilibrium constants based on concentrations and partial pressures, respectively. For the hypothetical system $2A(g) + B(g) \leftrightarrow 3C(g)$,
   $K_c = \frac{[C]^3}{[A]^2\,[B]}$ and $KP = \frac{(P_C)^3}{(P_A)^2\,(P_B)}$
6. According to Le Chatelier's Principle, when some change is made to an equilibrium system, the system reacts to partially counteract the imposed change and re-establish equilibrium.
7. When a *catalyst* is added to an equilibrium system, the rates of the forward and the reverse reactions increase to the same extent.
8. The concentrations of the reactants and the products remain unchanged when a catalyst is added to an equilibrium system.
9. In the presence of a catalyst, the value of the equilibrium constant, K_c, remains the same as in the initial equilibrium system.
10. When an inert gas like helium is introduced at constant volume, the total pressure of the system increases, but does not change the partial pressures (or the concentrations) of the reactants. Therefore, it has no effect on the system.
11. In a heterogeneous equilibrium system the equilibrium constant, K_c, does not include substances in the solid state.
12. There is no effect on the addition or removal of some solid reactant or product to a heterogeneous equilibrium system; LCP is not applicable in this case.

Figure 1 Major propositional content knowledge statements defining instruction on chemical equilibrium

An instructional program was developed by the chemistry teacher involved in the study in consultation with the first author based on the 12 propositional content knowledge statements (PCKSs) that were formulated by the authors (see Figure 1). The instructional program is found in Appendix 1. The lessons that were conducted over a period of four weeks (involving 2 hours 40 mins per week of instruction) addressed all the propositional content knowledge statements that were previously identified.

The instruction incorporated several instructional strategies including demonstrations of various chemical equilibrium reactions involving animations using the CD provided by the ministry (Malaysian Examinations Council, 2003), group discussions, individual presentations and problem solving. Students were able to follow the course of reactions and interpret the changes by plotting suitable graphs. In addition, the teacher used several demonstrations and provided opportunities for the students to conduct laboratory experiments.

During the lessons the teacher explained with notes on the blackboard and the students normally wrote their own notes. There was no specific textbook prescribed by the ministry but they were given a number of references (Loh & Sivaneson, 2004; Russo & Silver, 2010; Tan, 2012).

Instruments

In order to probe for a deeper understanding of students' views about chemical equilibrium we used two conceptual tests – the *Chemical Equilibrium Conceptual Test 1* (*CECT-1*) consisting of nine two-tier multiple-choice items that are convenient to administer and mark, and the *Chemical Equilibrium Conceptual Test 2* (*CECT-2*) with four short structured questions. The focus of the two instruments was to probe students' understanding of chemical equilibrium concepts more deeply than in previous studies. The instruments were administered one after the other as posttests after instruction. Students were given a total of 1½ hours to respond to the questions in both instruments. A pretest was not considered necessary because the students had not been instructed in chemical equilibrium during their previous years of secondary science studies.

The *CECT-1* consisted of nine two-tier multiple-choice items eight of which were adapted from previously developed items that have been documented in the research literature (Özmen, 2008; Cheung, 2009). Seven of the items were adapted from Özmen (2008) who was not specific about the conditions under which equilibrium was established; the modified items addressed this issue. One item, a multiple-choice item that was used by Cheung (2009) was adapted as a two-tier multiple-choice item by the authors. The ninth item was developed by the authors. These nine items provided an additional reference source to researchers, thus contributing to the existing research literature.

Two-tier multiple-choice items were used in the *CECT-1* because this was a convenient method for understanding students' reasoning for the choice of particular content responses (Treagust, 1988, 2006), as opposed to interviewing students which could be very time-consuming. Students had to select a content response from the first tier and subsequently select a reason response from the second tier to explain their choice in the first tier. These instructions were stated in the *CECT-1* and were also explained before students attempted the questions. They were given sufficient time to reflect on their choices and make changes if necessary. The *CECT-1* was a paper-and-pencil test that was convenient to administer and easy to mark. Correct responses to both tiers of each item were scored '1' while incorrect responses to both or either of the tiers were scored '0'.

The nine items in the *CECT-1* were related to (1) establishing equilibrium (Item 1), (2) adding more of a reactant to an aqueous equilibrium system (Item 2), (3) using LCP involving changes in the temperature of a gaseous system (Item 3), (4) the value of the equilibrium constant, K_{eq} (Items 4 & 5), (5) the effect of a catalyst on a system at equilibrium (Items 6 & 7), (6) a heterogeneous equilibrium system (Item 8) and (7) adding an inert gas to a gaseous equilibrium system (Item 9). The *CECT-1* had a Cronbach's alpha reliability coefficient of 0.63 that is above the threshold value of 0.50 for two-tier items (Nunally and Bernstein 1994).

The second instrument, the *CECT-2*, consisted of four structured questions developed by the authors to demonstrate issues identified in the literature and are presented in the next section. These items probed deep understanding of several concepts like (1) application of LCP, (2) computing the equilibrium constant, K_c, (3) variation of K_c with temperature, and (4) comparing equilibrium constants, K_c and K_p.

It was not possible to compute the Cronbach's alpha reliability coefficient for this instrument as only four structured items were used requiring short answers for responses. The *CECT-1* and *CECT-2* will add to the resources that are available in the extant research literature for use in further studies by other researchers.

Rationale for use of items in the CECT-1

Most of the items, except for Items 6 and 9, were adapted from a study conducted by Özmen (2008). All these seven items were not specific about the conditions under which the chemical equilibrium systems were established. In addition, the responses in one or both tiers needed to be worded more explicitly.

Item 1 about the gaseous equilibrium system, $PCl_5(g) \leftrightarrow PCl_3(g) + Cl_2(g)$, was adapted by specifying that the gaseous equilibrium systems was in a *closed reaction vessel*. In addition, the wording of two of the four reasons in the second tier was improved.

Item 2 about the equilibrium system, $2CrO_4^{2-}(aq) + 2H^+(aq) \leftrightarrow Cr_2O_7^{2-}(aq) + H_2O(l)$, three of the four reasons were reworded.

The original Item 3 was not specific about the conditions of the exothermic gaseous equilibrium system $N_2(g) + 3H_2(g) \leftrightarrow 2NH_3(g)$. This ítem was modified to indícate that the temperature of the system was increased at *constant pressure.* All four reasons were also modified.

Again, the original Item 4 did not specify the conditions of the exothermic gaseous equilibrium system $4NH_3(g) + 5O_2(g) \leftrightarrow 4NO(g) + 6H_2O(g)$. So, this item was modified to indicate that the temperature of the system was increased at *constant pressure.* Three of the four reasons were also modified.

The original Item 5 did not specify that the gaseous equilibrium system was in a *closed container* to suggest that the volume of the equilibrium system was constant. One of the four reasons was also modified for clarity.

Item 6 about the effect of adding a catalyst to the gaseous equilibrium system $2SO_2(g) + O_2(g) \leftrightarrow 2SO_3(g)$, was developed by the authors.

The fact that the gaseous equilibrium system $2CO(g) + O_2(g) \leftrightarrow 2CO_2(g)$ was in a *closed container* was not specified in the original item So, Item 7 in the *CECT-1* clarified that the gaseous system was at constant volume. In addition, two of the four reasons were improved.

The original Item 8 did not specify that the heterogeneous equilibrium system $CaCO_3(s) \leftrightarrow CaO(s) + CO_2(g)$, was in a *closed container*. In addition to specifying this condition, all the responses in both tiers were modified for clarity.

Item 9 about the effect of adding an inert gas to the gaseous equilibrium system $CO(g) + 2H_2(g) \leftrightarrow CH_3OH(g)$, was adapted from Cheung (2009). However, his item was a multiple-choice item. So, the authors developed the second tier of reasons.

Rationale for use of items in the CECT-2

The four short-answer items in this instrument were developed by the authors to probe more deeply students' understanding about basic principles related to chemical equilibrium.

RESULTS AND DISCUSSION

Analysis of responses to the CECT-1

A summary of students' responses to each of the items is discussed below. As the study investigated students' understanding of equilibrium concepts, the percentage of both the students' correct responses to each item as well as their major alternative conceptions have been provided. For convenience, alternative conceptions displayed by at least 10% of the students were considered. Using a figure greater than 10% could result in overlooking some major alternative conceptions.

Item 1: Students were required to predict the concentrations of the products when 0.30 mol of PCl_5 reached equilibrium in a closed container at constant temperature in the equilibrium system: $PCl_5(g) \leftrightarrow PCl_3(g) + Cl_2(g)$. Assuming that the container had a volume of 1L, only 31 students (55.4%) correctly suggested that the concentrations of the products were each less than 0.3 $molL^{-1}$ because only some of the PCl_5 had decomposed to establish equilibrium. This answer reflects understanding that the reaction (decomposition of PCl_5) does not go to completion.

Despite the relatively low percentage of students who answered this item correctly, only eight students (14.3%) held the highest alternative conception that when a certain amount of $PCl_5(g)$ reaches equilibrium in a closed container, the concentrations of all the products in the $PCl_5(g) \leftrightarrow PCl_3(g) + Cl_2(g)$ reaction mixture remain the same as that of the $PCl_5(g)$ because the concentrations of all the species are equal at equilibrium.

Item 2: This item involved an orange solution of 0.5 M $Na_2Cr_2O_7$ in which the following equilibrium system was established: $2CrO_4^{2-}(aq) + 2H^+(aq) \leftrightarrow Cr_2O_7^{2-}(aq) + H_2O(l)$. Students had to predict the change when 10 mL of orange 0.5 M $Na_2Cr_2O_7$ solution was added. Only seven students provided the correct response that the colour of the solution remained unchanged because there was no change in the concentration of any species. However, 18 students (32.1%), using LCP, displayed the misunderstanding that the solution turned yellow (due to the formation of more $CrO_4^{2-}(aq)$ ions) to counter the increased amount of $Cr_2O_7^{2-}(aq)$ ions.

Item 3: In predicting the effect of increasing the temperature on the ammonia equilibrium, $N_2(g) + 3H_2(g) \leftrightarrow 2NH_3(g)$, keeping the pressure constant, only 39 students (69.6%) correctly applied LCP to predict the effect of increasing the temperature at constant pressure. Of the remaining students, only nine (16.1%) held the alternative conception that the equilibrium shifts to the right when the temperature is increased favouring the formation of more ammonia.

Item 4: Students were required to predict the effect of increasing the temperature (at constant pressure) on the equilibrium constant for the oxidation of ammonia in the first step of the Ostwald process for the synthesis of nitric acid: $4NH_3(g) + 5O_2(g) \leftrightarrow 4NO(g) + 6H_2O(g)$, $\Delta H = -905.6$ kJ/mole. Only 23 students (41.1%) provided the correct response that as the forward reaction is exothermic, the reverse reaction is favoured when the temperature is increased. The increased rate constant for the reverse reaction results in a decrease in the equilibrium constant. However, 14 students (25%) held the misconception that the equilibrium constant increased with temperature.

Item 5: Item 5 involved the effect of changing the concentrations of the reactants in the hydrogen iodide equilibrium, $H_2(g) + I_2(g) \leftrightarrow 2HI(g)$, in a closed container at constant temperature. Only 33 students (58.9%) correctly suggested that the equilibrium constant remained unchanged despite changes in the concentrations of the reactants (or products). Of the remaining students, 11 students (19.6%) held the alternative conception that the equilibrium constant increased as the concentrations of the different species changed.

Item 6: This item involved predicting the effect of adding a catalyst to the gaseous sulfur trioxide equilibrium system, $2SO_2(g) + O_2(g) \leftrightarrow 2SO_3(g)$, $\Delta H = -197.78$ kJ/mole. A total of 34 students (60.7%) correctly suggested that the addition of a catalyst had no effect on the equilibrium constant as the catalyst lowers the activation energy of both the forward and reverse reactions to the same extent. However, eight students (14.3%) wrongly believed that the rate of the forward reaction rate increased although the catalyst acts by lowering the activation energy for both the forward and reverse reactions by the same amount.

Item 7: Students were required to predict the effect of adding a catalyst on the concentration of carbon dioxide in the equilibrium system involving the oxidation of carbon monoxide: $2CO(g) + O_2(g) \leftrightarrow 2CO_2(g)$, $\Delta H = -556.0$ kJ/mole. The inability of a catalyst to affect the concentrations of the components in the equilibrium mixture was indicated by 36 students (64.3%). No major alternative conception (equal to or greater than 10%) was identified for students' understanding of this item.

Item 8: This item involved the heterogeneous equilibrium of the decomposition of solid calcium carbonate in a closed container: $CaCO_3(s) \leftrightarrow CaO(s) + CO_2(g)$. Only 14 students (25%) correctly suggested that the removal of some solid calcium carbonate did not disturb the equilibrium. However, 26 students (46.4%) incorrectly applied LCP to suggest that more calcium carbonate was produced because $CaO(s)$ and $CO_2(g)$ had reacted to replace the calcium carbonate that was removed.

Item 9: Item 9 involved adding some argon to the equilibrium system $CO(g) + 2H_2(g) \leftrightarrow CH_3OH(g)$ at constant volume. Only 15 students (26.8%) correctly suggested that no change occurred because the partial pressures of the gases remained unchanged. However, 29 students (51.8%) suggested that there was no change merely because argon was not involved in the reaction.

As seen in Table 1, less than 50% of students provided correct responses to four of the nine items (Item numbers 2, 4, 8 and 9).

A summary of students' major alternative conceptions is provided in Table 2.

Further analysis showed that the achieved overall scores after instruction ranged from 0 to 8 with a mean score of 4.14 (out of a maximum of 9) (see Figure 2).

Table 1. Frequency of correct responses to both tiers of each item in the *CECT-1* (N = 56)

Item no.	Combined tiers frequency	Item no.	Combined tiers frequency
1	31 (55.4)	6	34 (60.7)
2	7 (12.5)	7	36 (64.3)
3	39 (69.6)	8	14 (25.0)
4	23 (41.1)	9	15 (26.8)
5	33 (58.9)		

Note: percentages are in parentheses

Table 2. Major alternative conceptions held by students about items in the *CECT-1* (N = 56)

Item no.	Alternative conceptions	Percentage of students
1.	When a certain amount of $PCl_5(g)$ reaches equilibrium in a closed container, the concentrations of all the products in the $PCl_5(g) \leftrightarrow PCl_3(g) + Cl_2(g)$ reaction mixture remain the same as that of the $PCl_5(g)$ because the concentrations of all the species are equal at equilibrium.	14.3
2.	If more 0.5M solution of sodium dichromate is added to the original orange 0.5M solution of sodium dichromate, the solution turns yellow (due to the formation of more $CrO_4^{2-}(aq)$ ions) to counter the increased amount of $Cr_2O_7^{2-}(aq)$ ions.	32.1
3.	The exothermic $N_2(g) + 3H_2(g) \leftrightarrow 2NH_3(g)$ equilibrium shifts to the right when the temperature is increased, favouring the formation of more ammonia, when the temperature of the system is instantaneously increased.	16.1
4.	The equilibrium constant of the exothermic $4NH_3(g) + 5O_2(g) \leftrightarrow 4NO(g) + 6H_2O(g)$ equilibrium system increases with temperature as the rate constant of the reaction always increases with temperature.	25.0
5.	The equilibrium constant of the $H_2(g) + I_2(g) \leftrightarrow 2HI(g)$ system increases as the concentrations of the different species changes.	19.6
6.	The rate of the forward reaction of the $2SO_2(g) + O_2(g) \leftrightarrow 2SO_3(g)$ equilibrium system increases when a catalyst is added although the catalyst acts by lowering the activation energy for both the forward and reverse reactions by the same amount.	14.3
8.	More calcium carbonate is produced after some solid $CaCO_3$ is removed from the $CO(g) + 2H_2(g) \leftrightarrow CH_3OH(g)$ equilibrium mixture in a closed container because $CaO(s)$ and $CO_2(g)$ react to replace the calcium carbonate that is removed.	46.4

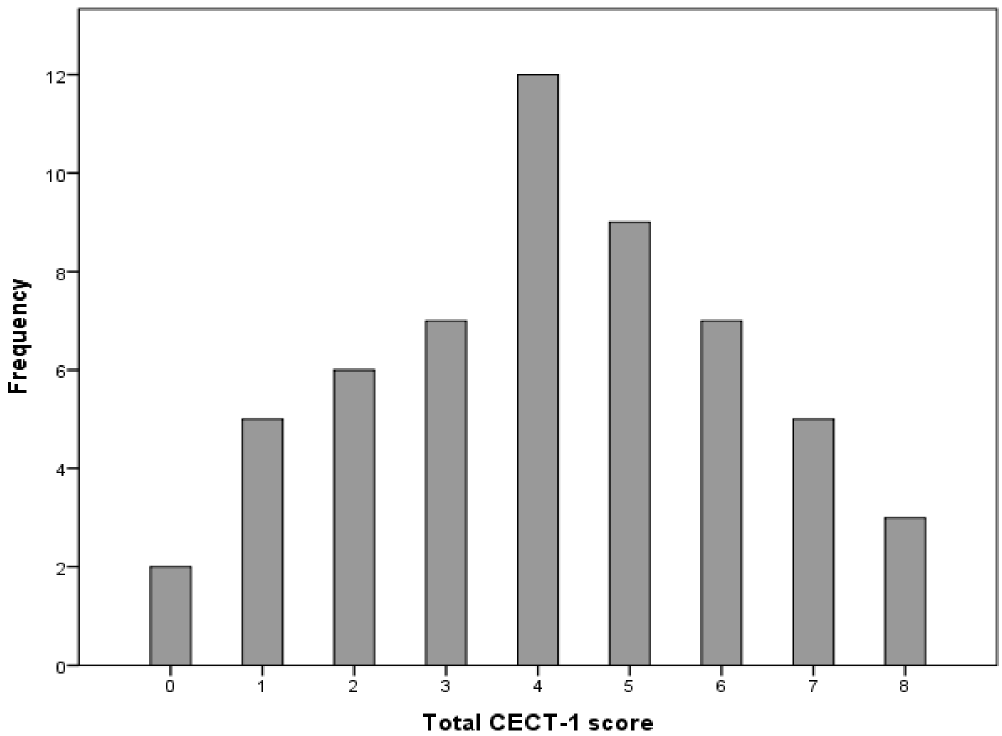


Figure 2. Distribution of total *CECT-1* scores (N = 56)

Analysis of responses to the CECT-2

Students' responses to the four items after instruction (in Figures 3 – 6) are discussed below. Unfortunately, the responses from students were very lacking as in many instances there were no responses at all from several students. As a result, it was not possible to provide a list of major alternative conceptions that were held by the students.

Item 1 (Le Chatelier's Principle)

The equilibrium system represented by the equation is established in a container below fitted with a piston (Diagram A).

Diagram A

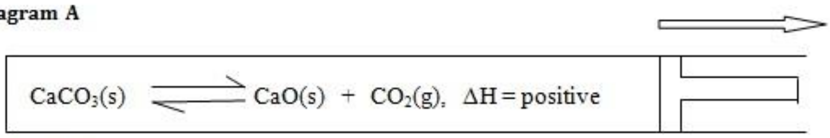
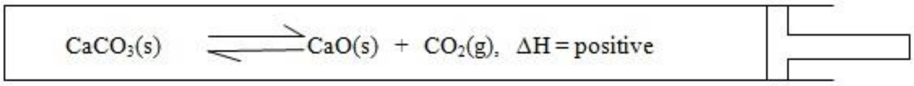


Diagram B



1(a) The piston is then moved in the direction shown to the position in diagram B.
(i) Immediately after going from A to B at constant temperature, what happens?
(ii) How would the system respond to this change?
(iii) In which direction would the system shift in order to re-establish equilibrium?

1(b) The temperature of the system is increased at constant pressure in diagram A.
(i) Which reaction would this change favour?
(ii) Why do you say so?
(iii) In which direction would the system shift in order to re-establish equilibrium?

1(c) What is the net change in the number of moles of the products when equilibrium is re-established under both the conditions in (a) and (b)?

Figure 3. Item 1 in the *CECT-2*

Section 1a (3 marks): Only 10 students (17.9%) correctly suggested in their responses that when the volume of the equilibrium system is increased at constant temperature, the pressure of the system would decrease (or the volume of CO_2 would increase) by more $CaCO_3$ decomposing; equilibrium is thus reestablished by the system shifting from left to right.

Most of the remaining 46 students suggested that the pressure of the system would increase because more $CaCO_3$ decomposed to form more CO_2.

Section 1b (3 marks): Given that ΔH of the forward reaction was positive, 47 students (83.9%) correctly suggested that when the temperature of the system was increased at constant pressure, the equilibrium would shift from left to right in favour of the forward reaction. All these students explained that as this reaction was endothermic, heat energy would be absorbed in response to b (ii) of the question. The remaining nine students were unable to provide an answer.

Section 1c (1 mark): A total of 39 students (69.6%) correctly suggested that under conditions of constant temperature and pressure stated in sections (a) and (b), the net change was an increase in the amount of products when equilibrium was re-established. The remaining 18 students did not respond to this part of the question.

The total scores for Item 1 ranged from 2 to the maximum of 7 that was achieved by only six students (10.7%).

For the esterification equilibrium system $P(l) + Q(l) \leftrightarrow R(l) + H_2O(l)$ the total scores ranged from 1 to 4; the maximum of 4 was achieved by 24 students (42.9%).

Item 2 (Equilibrium constant, K_c)

A mixture of
0.60 mol of a carboxylic acid, P,
0.50 mol of an alcohol, Q,
0.60 mol of the corresponding ester, R, and
0.40 mol of water,
in a total volume of V dm^3 is heated under reflux with a catalyst.
When equilibrium is established and on cooling the mixture to room temperature, only 0.40 mol of the carboxylic acid, P remained.
Using the steps provided below, deduce the equilibrium constant, K_c for the reaction:

$$P(l) + Q(l) \rightleftharpoons R(l) + H_2O(l).$$

(i) initial concentrations:
(ii) equilibrium concentrations:

(iii) Calculate the equilibrium constant, K_c for the reaction at room temperature.
(Use the steps provided below).
Write the expression for K_c:
Show concentrations in terms of moles and volume:
Deduce expression in terms of number of moles:
Hence, calculate K_c:

Figure 4. Item 2 in the *CECT-2*

Section (i) and (ii) (2 marks): A total of 36 students (64.3%) provided the correct initial and equilibrium concentrations for all the three species (P, Q and R) in the esterification equilibrium. The rest of the 20 students left this section blank.

Section (iii) (2 marks): Only 24 of the 36 students (66.7%) above provided the correct steps in the solution of K_c for the equilibrium system, arriving at an answer of 4.

Item 3 (Variation of the Equilibrium constant, K_c with temperature)

The equilibrium constant, K_c for the equilibrium system

$$N_2(g) + O_2(g) \rightleftharpoons 2NO(g),$$

was found to increase as the temperature increased.
(a) What does this tell us about the sign of ΔH for the forward reaction?
(i) the forward reaction is ....
(ii) as the temperature decreases, the .... is favoured.
(iii) the percentage of the product, NO ....as the temperature increases.
(b) Based on the chemical equation, we may also conclude that the percentage of the product, NO .... as the pressure of the system increases.
(c) Suggest a reason for your answer in (b).

Figure 5. Item 3 in the *CECT-2*

The total score for Item 3 ranged from 0 to 5; the maximum score was achieved by only one student (1.8%).

Section 3(a) (3 marks): Given that the equilibrium constant K_c for the equilibrium system increased with temperature, only 14 students (25%) correctly suggested that ΔH for the forward reaction was positive (i.e. the forward reaction was endothermic). These students then went on to suggest that the reverse reaction was favoured when the temperature was decreased; an increase in temperature therefore favoured the formation of more product, NO(g).

Sections 3(b) and 3(c) (1 mark each): Of the 14 students referred to above, only one student correctly suggested that the percentage of product, NO(g), remained unchanged when the pressure of the system was increased because there was no change in the number of moles of gaseous reactants and products in the equilibrium system. The remaining 22 students did not attempt this question.

Item 4 (Equilibrium constants, K_c and K_P)

For the gaseous system that is in equilibrium as shown below,

$$A(g) \rightleftharpoons 2B(g),$$

KP is given by the expression:

$$KP = \frac{(P_B)^2}{(P_A)}$$

$$= \frac{(nB\ /\ nA{+}B\ x\,P)^2}{(nA\ /\ nA{+}B\ x\,P)}$$

where, P_A and P_B are the partial pressures of A and B respectively,
n_A and n_B are the number of moles of A and B respectively, and
P is the total pressure of the system.

(a) Deduce K_c and K_P for the equilibrium system shown below:

$$3Fe(s) + 4H_2O(g) \rightleftharpoons Fe_3O_4(s) + 4H_2(g).$$

Expression for K_c:
In terms of moles:
Similarly for K_p:
(b) What do you notice about the values of K_c and K_p for this equilibrium reaction?
(c) Suggest a reason for your answer to (b).

Figure 6. Item 4 in the *CECT-2*

The total score for Item 4 ranged from 0 to 4 with none of the students achieving the maximum score of 6 for the hypothetical A(g) ↔ 2B(g) gaseous equilibrium system. A total of 16 students (28.6%) did not attempt this item.

Section 4(a) (4 marks): Only two students derived the correct expressions for K_c and K_p for the above hypothetical equilibrium system.

Sections 4(b) and 4(c) (1 mark each): Neither of the above could deduce that the two values were equal because there was no change in the number of gaseous reactants and products in the equilibrium system.

The distribution of the total scores achieved by students in the *CECT-2* after instruction ranged from 7 to 17 (Figure 7) with a mean score of 11.0 out of a maximum score of 22. Almost half the number of students (44.6%) scored less than 50% of the total marks.

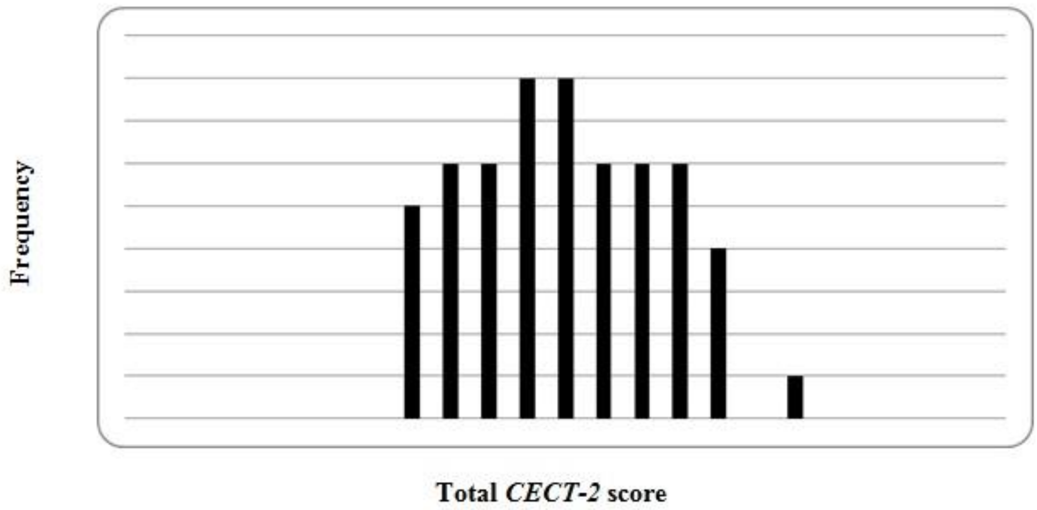


Figure 7. Distribution of total *CECT-2* scores (N = 56)

Students' responses to Item 1 in the *CECT-2* reinforced their limited understanding about the use of LCP in heterogeneous equilibrium systems that was identified in Item 9 of the *CECT-1*. Students generally experienced difficulty in computing the equilibrium constant, K_c, in Item 2 of the *CECT-2* with only 42.9% being successful, reinforcing lack of understanding that was demonstrated in Item 6 of the *CECT-1* when only 58.9% of students correctly suggested the correct value of K_{eq} for the $H_2(g) + I_2(g) \leftrightarrow 2HI(g)$ equilibrium system (see Table 2). Furthermore, none of the students were able to account for the similarity in vales of K_p and K_c for the heterogeneous equilibrium system $3Fe(s) + 4H_2O(g) \leftrightarrow Fe_3O_4(s) + 4H_2(g)$ in Item 4 of the *CECT-2*. While 41.1% of students correctly predicted the effect of increasing the temperate on the equilibrium constant for the $4NH_3(g) + 5O_2(g) \leftrightarrow 4NO(g) + 6H_2O(g)$, equilibrium system in Item 5 of the *CECT-1,* it was surprising that only 1.8% of students correctly answered Item 3 in the *CECT-2* regarding the variation of K_c with temperature for the $N_2(g) + O_2(g) \leftrightarrow 2NO(g)$ equilibrium system. The data obtained from the two instruments suggest that students were not consistent in their understanding of the chemical equilibrium concepts.

Analysis of students' responses showed that 0% to 42.9% of students scored the maximum possible marks for each item in the *CECT-2*. Achievement in all items was below 50%. Hence, it may be concluded that the level of understanding of chemical equilibrium concepts by the students was very limited with several students not attempting most of the questions.

CONCLUSION AND IMPLICATIONS

This research study has shown that the Year 12 Malaysian students possessed limited understanding of chemical equilibrium concepts. The two instruments that were used in this study reinforced some of the difficulties that were experienced by the students. For example, uncertainty about the use of LCP in heterogeneous equilibrium systems involving solids in particular, was evident in students' responses to items in both instruments. Similarly, students displayed limited ability when computing the equilibrium constants in items in both the instruments. The effect of temperature on the value of the equilibrium constant however, produced conflicting responses from the students, with 41.1% making correct predictions about a reaction in the first instrument while only 1.8% provided a correct response to another reaction in the second instrument. The students displayed other misunderstandings like the reversibility of chemical reactions leading to establishing a state of equilibrium, the effect of a catalyst or an inert gas on an equilibrium system. When using LCP in particular, it is important to be aware of the conditions under which the system concerned is being disturbed. The simplistic definition of LCP is likely to have caused serious errors in its application to equilibrium systems.

Especially in gaseous systems, it is important that the conditions of equilibrium are unequivocally stated. The Malaysian HSC syllabus includes the ammonia equilibrium with particular reference to the Haber process. Based on anecdotal evidence, the discussion of this equilibrium system is seldom referred to at constant volume (probably because it is assumed that students are aware that the industrial manufacture of ammonia occurs in a closed container); LCP is then conveniently used to predict the effects of changes in temperature and pressure.

It is important therefore, that high school students have a firm foundation about the basic concepts of chemical equilibrium in preparation for more advanced study at university. Educators in pre-service teaching programs need first to be made aware of students' misconceptions so that they are able to take appropriate measures to convey the correct ideas about chemical equilibrium concepts to their student teachers. At the same time, well-coordinated professional development courses need to be considered for in-service teachers. In regards to LCP in particular, changes in the contents of textbooks are necessary because these are a major source of information that teachers and students depend on. As Cheung (2009) suggests that re-education of "teacher educators, textbook writers and school teachers, on the inadequacy of LCP should be a high priority" (p. 518).

To enhance students' understanding of chemical equilibrium concepts, it may be necessary for high school chemistry teachers to make greater use of alternative instructional strategies like computer simulations (Khan, 2011) to solve conceptual problems; several of these computer simulations are freely available on the Internet. There is no certainty, however, that the use of computer simulations will necessarily benefit all students. Studies involving college students in the US, for example, have shown that depending on their prior knowledge, students use computer animations differently when solving problems. In particular, students with lower prior knowledge of chemistry concepts were more likely to use computer simulations to confirm their predictions about particular concepts than their peers with higher prior knowledge (Liu et al., 2008).

In particular, teachers will need to spend more time ensuring that students are better able to answer questions like the ones in the *CECT-2*. Answering questions of this nature will further enhance students' understanding of chemical equilibrium concepts.

Finally, the study needs to be extended to involve a larger sample of students of varying achieving abilities from a variety of schools in Malaysia. In addition, a comparative study involving similar grade students from several countries would be beneficial to better gauge the extent of the issues that have been identified in this study.

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APPENDIX

Chemical equilibrium instructional program

Week no.	Outline of lessons
1	**Topic:** Establishing equilibrium **Objectives of the lesson:** 1. To describe a reversible reaction and dynamic equilibrium. 2. To define the law of mass action and derive the equilibrium constant (K_c and K_p) for the general equation aA + bB ↔ cC + dD **Stage I. Introduction** 1. Comparison of reversible reactions (e.g., Haber process) and non-reversible reactions (e.g., combustion process) which students had learned in Form 5. 2. For the students to be able to distinguish between reversible and non-reversible reactions, the animation provided in the CD from the ministry was used. For the non-reversible reaction, the animation of the reaction between 6 moles of magnesium and 125cm^3 HCl was used. The students had to follow the instructions and click the relevant icons to view the particles present. For the reversible reaction, the animation of the reaction between H_2 and I_2 to form HI was used. 3. A reversible reaction will never go to completion to produce 100% of product. 4. Following questions were asked: (i). Describe the reaction in the Haber process and combustion reactions. (ii). Write the chemical equation for both the processes. (iii). Include the states of the starting materials and the products in the equation. 5. Identify the differences between both the processes. 6. The students were required to make note of the changes in the concentrations of product and reactants (in the HI equilibrium in the CD) with time. The concept of chemical equilibrium was introduced when the concentrations remained unchanged after a certain time. 7. The students were then asked to draw a graph of concentration vs. time by clicking the graph icon in the CD. The concepts were further illustrated by demonstrating the reaction of H_2 and I_2 to form HI using animations. 8. Students are shown a graph of concentration versus time for the reaction in the Haber process to see how the concentration of both reactants and product change with time and later remain unchanged. 9. The concept of dynamic equilibrium was described using the graph that shows rate versus time. **Stage II. Equilibrium Law (Law of Mass Action)** 1. Use a general equation aA + bB ↔ cC + dD to explain how to derive equilibrium constant (K_c and K_p). 2. Using the same general equation, the law of mass action was explained to the students. **Stage III. Class Discussion** 1. Students discussed in groups the heating of solid calcium carbonate in an open container and in a closed container. Then the teacher explained to the class which reactions were reversible and non-reversible. 2. The teacher guided the discussions and the presentations. **Stage IV. Application** The students were requested to solve following question.

	The equilibrium constant, K_p, for the reaction $CO(g) + 2H_2(g) \leftrightarrow CH_3OH(g)$ is 6.7 × 10^{-5} atm^{-2} at 320 °C. If CO at 5.0 × 10^{-2} atm, H_2 at8.0 × 10^{-1} atm, and CH_3OH at 7.0 × 10^{-5} atm are mixed together at 320 °C, determine: (i) Whether the system is in a state of equilibrium; (ii) The direction of the net reaction if the system is not in equilibrium.
2	**Topic:** Homogeneous and heterogeneous equilibrium systems **Objectives of the lesson** 1. Deduce expressions for the equilibrium constant in terms of concentration K_c, and partial pressures, K_p, for homogeneous and heterogeneous systems. 2. Calculate the values of the equilibrium constants in terms of concentration or partial pressures from given data. 3. Calculate the quantities of substances present at equilibrium from given data. **Stage I. Introduction** 1. Students were shown an example each of heterogeneous and homogenous reactions and the corresponding equations. 2. For a heterogeneous reaction, students were reminded of the heating of calcium carbonate which students had learned in the previous lesson. 3. The teacher explained how to calculate K_p and K_c 4. Using the general equation $aA + bB \leftrightarrow cC + dD$, the teacher explained how to derive an expression to relate K_p and K_c for the general gas-phase equation and showed when was K_p=K_c. For this purpose the ideal gas equation (pV= nRT) was used. **Stage II** 1. Students were provided with the esterification equation and were requested to calculate the value of K_c. 2. Students were provided with the equation of the reaction of hydrogen and iodine gases and requested to calculate K_p . **Stage III. Class Discussion** Students worked in groups and presented the answers for the questions raised in stage II. **Stage IV. Application** Students were requested to solve the following problems. Q1: Consider the two equilibria shown below. Reaction I: $X_2(g) + 2Y_2(g) \leftrightarrow 2XY_2(g)$ Reaction II: $XY_2(g) \rightleftharpoons \frac{1}{2}X_2(g) + Y_2(g)$ The numerical value of K_c for reaction I is 4. What is the equilibrium constant, $K_c{}'$, for reaction II? Give the units for$K_c{}'$. Q2: The equilibrium constant for the reaction $H_2(g) + I_2(g) \leftrightarrow 2HI(g)$ is 60 at 450 °C. Calculate the number of moles of HI in equilibrium with 2.0 mol of H_2 and 0.30 mol I_2 at 450 °C. Q3: Solid ammonium chloride decomposes on heating as shown below: $NH_4Cl(s) \leftrightarrow NH_3(g) + HCl(g)$. In an experiment carried out at 400 °C, the equilibrium mixture was found to contain 32.6 g of $NH_4Cl(s)$, and $NH_3(g)$ and $HCl(g)$ with partial pressures of 2.38 atm and 5.04 atm, respectively. Calculate the equilibrium constant, K_p, at 400 °C.
3	**Topic:** Factors affecting chemical equilibria: Le Chatelier's Principle **Objectives of the lesson** 1. State the Le Chatelier's principle and use it to discuss the effects of catalyst, changes in concentration, pressure or temperature on a system at equilibrium in the following examples:

	(i) synthesis of hydrogen iodide, (ii) dissociation of dinitrogen tetroxide, (iii) hydrolysis of simple esters, (iv) the Contact process, (v) the Haber process, (vi) the Ostwald process. Discuss the effect on equilibria reactions of (i) Concentration (in the synthesis of hydrogen iodide), (ii) Pressure (in the Haber process), (iii) Adding a noble gas at constant volume, (iv) Adding a noble gas at constant pressure, (v) Catalyst (the hydrolysis of simple esters). **Stage I. Introduction** 1. The students were requested to recall the Haber process that they had learned earlier. 2. The teacher defined Le Chatelier's Principle and explained its use in the reaction in the Haber process. **Stage II** The teacher used the following reactions to explain the changes/effect on the equilibrium of: (i) Concentration (in the synthesis of hydrogen iodide), (ii) Pressure (in the Haber process), (iii) Adding a noble gas at constant volume, (iv) Adding a noble gas at constant pressure, (v) Catalyst (the hydrolysis of simple esters). **Stage III. Class Discussion** Using the Haber and Contact processes as examples, the students were requested to work in groups and present the effects of temperature and pressure increases on the equilibrium for both the processes. **Stage IV. Application** Students were provided with a worksheet with questions based on Le Chatelier's Principle requiring them to state what was observed when the equilibrium states of the systems were disturbed.
4	**Topic:** Equilibrium constant in terms of partial pressures and concentrations **Objective of the lesson** Explain the effect of temperature on the equilibrium constant in the equation $$\ln K = -\frac{\Delta H}{RT} + c$$ (i) Discuss the effect of temperature on K_c and K_p for endothermic and exothermic reactions (using Le Chatelier's principle) (ii) Use the van't Hoff equation to calculate the change of K_c with temperature. $$\ln K = -\frac{\Delta H}{RT} + c$$ (iii) Use the van't Hoff equation to calculate ΔH based on data given. $$\ln \frac{K_1}{K_2} = -\frac{\Delta H}{R}\left(\frac{1}{T_1} - \frac{1}{T_2}\right)$$ **Stage I. Introduction** 1. The students were told that the effect of temperature on the equilibrium constant is given by $$\ln K = -\frac{\Delta H}{RT} + c$$ 2. The students were shown graphs of ln K against 1/T for both endothermic and exothermic reactions and explanations were given. **Stage II**

	The students were requested to calculate the equilibrium constant of a reaction at 1273 K when the equilibrium constant of the same reaction at 1073K was given as 2.00 and ΔH= +16.7kJ. **Stage III** The students were requested to solve in groups, similar examples given by the teacher. **Stage IV Application** For the disassociation of N_2O_4(g) to NO_2 (g), the value of K_p and temperature were given. Students were required to plot a graph and determine the enthalpy change for the reaction.