



THE SCHOOL FOR EXCELLENCE
UNIT 4 CHEMISTRY 2008
AREA OF STUDY 1 TEST – SOLUTIONS

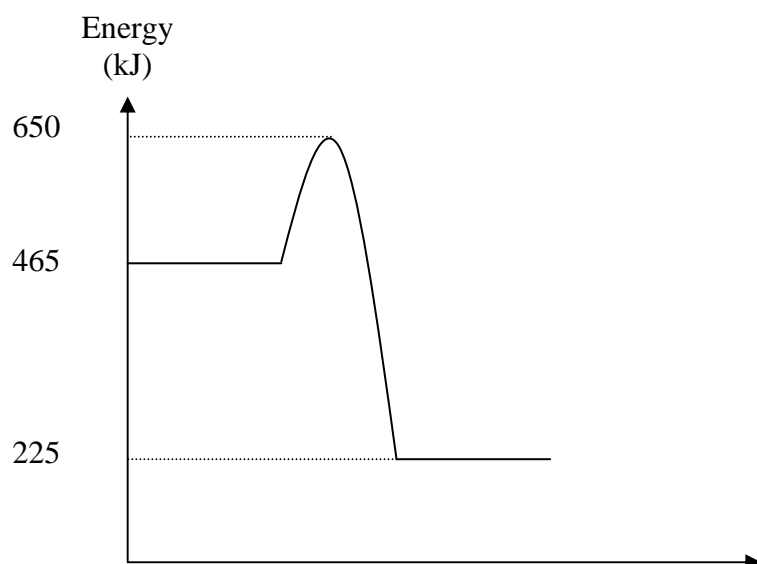
SECTION A – MULTIPLE CHOICE QUESTIONS

- | | |
|--------------------|-------------|
| QUESTION 1 | Answer is B |
| QUESTION 2 | Answer is C |
| QUESTION 3 | Answer is D |
| QUESTION 4 | Answer is C |
| QUESTION 5 | Answer is D |
| QUESTION 6 | Answer is D |
| QUESTION 7 | Answer is B |
| QUESTION 8 | Answer is D |
| QUESTION 9 | Answer is C |
| QUESTION 10 | Answer is A |
| QUESTION 11 | Answer is A |
| QUESTION 12 | Answer is C |
| QUESTION 13 | Answer is C |
| QUESTION 14 | Answer is D |
| QUESTION 15 | Answer is C |
| QUESTION 16 | Answer is C |
| QUESTION 17 | Answer is A |
| QUESTION 18 | Answer is C |
| QUESTION 19 | Answer is C |
| QUESTION 20 | Answer is D |

SECTION B – SHORT ANSWER QUESTIONS

QUESTION 1

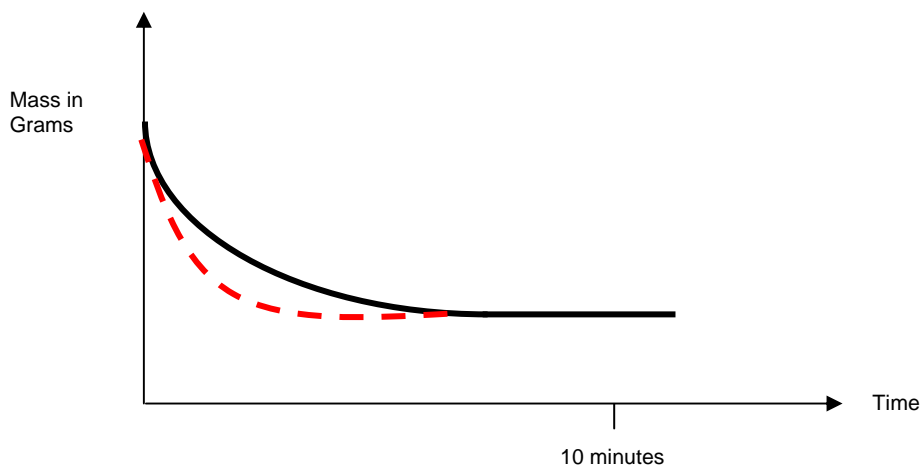
- a. (i) -480 kJ/mol
- (ii) 370 kJ/mol
- (iii) Catalysts affect the rate of the back reaction to the same extent as the forward reaction. i.e. the back reaction rate would increase.
- b. (i) 850 kJ/mol
- (ii) The bonds in the products are stronger as it takes more energy to convert the products to the activated complex than it does to convert the reactants to the activated complex.
- c. Halve every marked quantity.



QUESTION 2

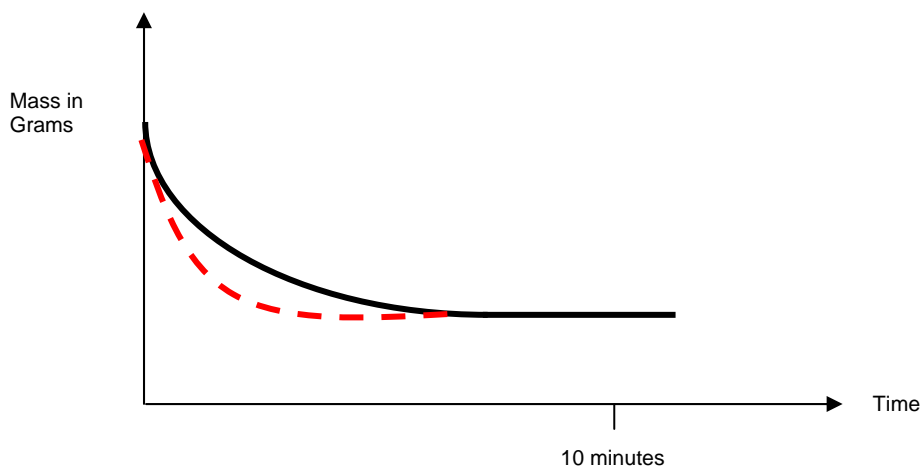
Note: As vessel is open, the gaseous products will escape, hence the mass of the vessel and its contents will decrease.

- a. The calcium carbonate is ground into a finer powder.



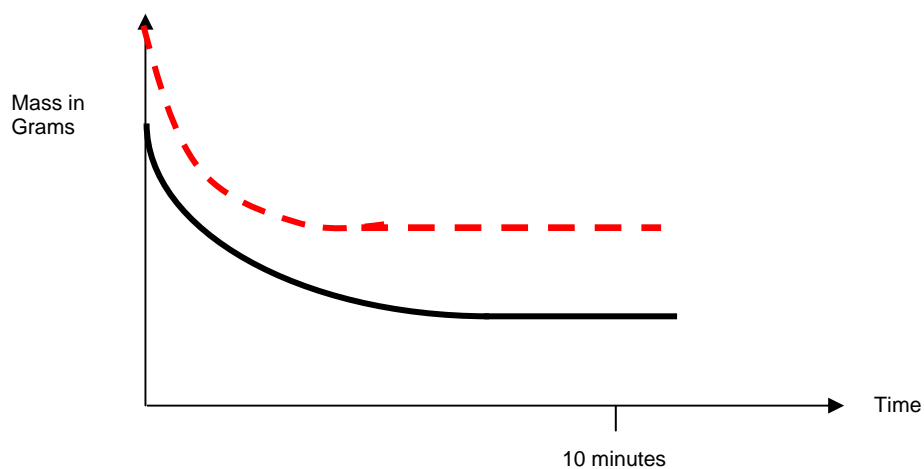
The reaction rate increases, but the initial and final masses are the same.

- b. The concentration of hydrochloric acid is increased to 1.0 M keeping volume and temperature constant.



The reaction rate increases, but the initial and final masses are the same.

- c. 400 ml of 0.50 M hydrochloric acid is used instead of 200 ml at constant temperature.



The extra 200 ml of hydrochloric acid adds to the initial mass. The initial mass will therefore be higher.

As the hydrochloric acid is in excess, more acid will be left over than in the previous examples. Therefore, the final mass will be higher.

The initial reaction rates are the same.

The reaction rates at all other times is higher as the number of acid particles remaining at each unit time is greater, meaning that the remaining concentration is higher.

QUESTION 3

a.
$$K = \frac{[SO_{3(g)}]^2}{[SO_{2(g)}]^2 [O_{2(g)}]} M^{-1}$$

b.



Initial Mole: 2.000 2.000 0.000

Final Mole: 0.150

Therefore: $2.000 - 0.150 = 1.850$ mole of SO_2 reacted.

Mole Ratio: 2 1 2

Amount Reacted: 1.850 $\frac{1.850}{2}$ 1.850

Amount at Eq: 0.150 $2.000 - \frac{1.850}{2} = 1.075$ $0 + 1.850 = 1.850$

Conc at Eq: $0.0150 M$ $0.1075 M$ $0.1850 M$ $\left(C = \frac{n}{V} = \frac{n}{10} \right)$

$$K = \frac{[SO_{3(g)}]^2}{[SO_{2(g)}]^2 [O_{2(g)}]} = \frac{(0.1850)^2}{(0.0150)^2 (0.1075)} = 1.41 \times 10^3$$

c.
$$K_{reverse} = \frac{1}{K_{forward}} = \frac{1}{1.41 \times 10^3} = 7.09 \times 10^{-4}$$

- d. Perform the reaction at different temperatures using the same amounts of each reactant. Then measure the amount of one of the species eg. SO_2 once the system reaches equilibrium and determine the value of K at that temperature.

By determining whether K increases or decreases as the reaction temperature increases, we can determine whether the forward reaction is exothermic or endothermic.

If K increases as temperature increases, the forward reaction is endothermic.

If K decreases as temperature increases, the forward reaction is exothermic.

Note: You cannot determine the nature of the reaction by say adding a reactant and driving the reaction forward, as this will result in a consequential temperature change, which will cause the position of equilibrium to shift, as well as alter the K value. It will be impossible to determine which variable was responsible for the observed final changes.

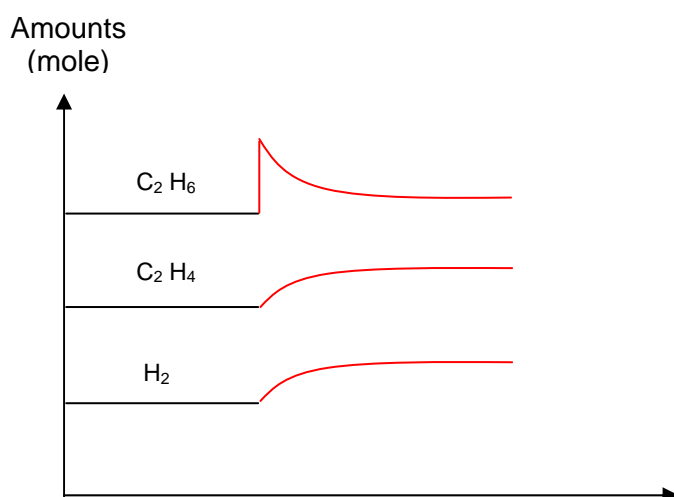
QUESTION 4

- a. (i) By decreasing the volume of the container, the total pressure within the system increases.

The system will partly compensate for this change by favouring the reaction that will decrease the pressure i.e. the reaction that will decrease the number of gaseous particles.

A net back reaction is favoured i.e. the position of equilibrium shifts to the left.

- (ii) When the volume of the container decreases, the concentration of ethene increases. Although the amount of ethene then decreases, the concentration never returns to its previous value. The final equilibrium concentration of ethene would therefore be higher than in the previous equilibrium mixture.
- (iii) The final forward and back reaction rates will be equal when the system returns to equilibrium, however, these rates will be higher than that rates in the initial equilibrium mixture. This increase in rates is due to the fact that all concentrations increased when the volume of the container was decreased. As the concentration of particles increases, the probability of collisions increase, hence reaction rates increase.
- b. Inert gases do not change the partial pressures of the gaseous reactants and products and hence their concentrations. Therefore, the system stays in equilibrium, and there is no change in the position of equilibrium.
- c. The system will compensate for the change by favouring the reaction that would decrease the concentration of ethane. A net forward reaction will occur. i.e. The position of equilibrium will shift to the right.



QUESTION 5

a. $K_w = [H_3O^+_{(aq)}][OH^-_{(aq)}]$

In pure water: $[H_3O^+] = [OH^-]$

$$0.68 \times 10^{-14} = [H_3O^+]^2$$

$$\therefore [H_3O^+] = \sqrt{0.68 \times 10^{-14}} = 8.25 \times 10^{-8}$$

$$pH = -\log_{10} 8.25 \times 10^{-8} = 7.08$$

b. (i) $[OH^-] = [NaOH] = 0.0040 \text{ M}$

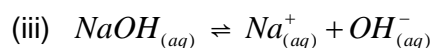
$$[H^+] = \frac{1 \times 10^{-14}}{0.0040} = 2.5 \times 10^{-12} \text{ M}$$

$$pH = -\log_{10}(2.5 \times 10^{-12}) = 11.6$$

(ii) $[OH^-] = [NaOH] = 0.0040 \text{ M}$

$$[H^+] = \frac{0.68 \times 10^{-14}}{0.0040} = 1.7 \times 10^{-14} \text{ M}$$

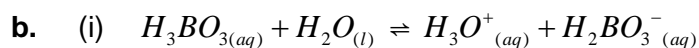
$$pH = -\log_{10}(1.7 \times 10^{-14}) = 13.8$$



As the temperature increases, the pH of the solution decreases, indicating that fewer hydroxide ions are being produced. This suggests that an increase in temperature favours a net back reaction. This would only occur if the forward reaction was exothermic.

QUESTION 6

a. The value of K_a for boric acid is very small, indicating that it is an extremely weak acid. Since the corrosive nature of an acid depends on the $[H_3O^+]$, this would be very low in boric acid and a dilute solution could be safely used without causing harm.



$$K_a = \frac{[H_3O^+]^2}{[H_3BO_3]} = 7.3 \times 10^{-10}$$

$$\therefore [H_3O^+] = \sqrt{7.3 \times 10^{-10} \times 0.5} = 1.9 \times 10^{-5} M$$

$$pH = -\log_{10} 1.9 \times 10^{-5} = 4.7$$

(ii) The approximation can be applied because the K_a value is less than 1.0×10^{-4} .

$$(iii) \% \text{ Hydrolysis} = \frac{[\text{One Product}]}{[\text{Initial Reactant}]} = \frac{[H_3O^+]}{[H_3BO_3]} \times 100$$

$$\% = \frac{1.9 \times 10^{-5}}{0.50} \times 100 = 0.0038\%$$