



## EQUILIBRIUM SL PAPER 1A

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

## EQUILIBRIUM SL PAPER 1B

1. (a) The yield increases [1]. Le Chatelier's principle dictates the system will shift to the right to reduce the pressure, as there are 4 moles of gas on the left and only 2 on the right [1].

(b) According to the table, as temperature increases (e.g., 300K to 700K), the yield of ammonia drastically decreases (e.g., 98% to 16%) [1]. An increase in temperature favors the endothermic direction [1]. Therefore, the reverse reaction is endothermic, making the forward reaction exothermic [1].

(c) At 300 K, the rate of reaction is far too slow for industrial viability [1]. 700 K is a compromise temperature that provides a faster rate of production, even at the expense of maximum theoretical yield [1].

2. (a)  $K_c = [Y]^2 / [X]$  [1].

(b) Adding X increases the denominator, so  $Q_c$  immediately decreases ( $Q_c < K_c$ ) [1]. The system shifts to the right (towards products) to re-establish equilibrium [1].

## EQUILIBRIUM SL PAPER 2

1. (a) The system reaches dynamic equilibrium at  $t = 25$  s, as the concentrations stop changing [1].

(b)  $K_c$  is greater than 1 [1]. The graph shows the equilibrium concentration of the product is higher than the equilibrium concentration of the reactant, meaning the numerator is larger than the denominator in the  $K_c$  expression [1].

2. (a)  $K_c' = 1 / K_c = 1 / 280 = 0.00357$  [1].

3. (a) The solution turns orange [1]. Adding HCl increases the  $[H^+]$ . By Le Chatelier's principle, the equilibrium shifts to the right to consume the added  $H^+$  [1], producing more orange  $Cr_2O_7^{2-}$  [1].

(b) The solution turns yellow [1]. The  $OH^-$  from NaOH neutralizes and removes the  $H^+$  ions from the equilibrium [1]. The equilibrium shifts to the left to replace the lost  $H^+$ , forming more yellow  $CrO_4^{2-}$  [1].

4. (a) A catalyst increases the rates of both the forward and reverse reactions equally [1], allowing equilibrium to be reached faster without altering the final equilibrium position or concentrations [1].

5. (a)  $CO_2(aq) \rightleftharpoons CO_2(g)$  [2] (1 for species, 1 for correct states).

(b) Opening the bottle suddenly decreases the pressure of  $CO_2(g)$  above the liquid [1]. By Le Chatelier's principle, the equilibrium shifts to the right to produce more gas, causing rapid bubbling as dissolved gas escapes [1].