EQUILIBRIUM HL PAPER 1A

1. B
2. C
3. B
4. B
5. D

6. C
7. A
8. B
9. B
10. B

11. A
12. B
13. B
14. B
15. A

16. B
17. C
18. B
19. C
20. C

21. A
22. B
23. B
24. A
25. D

26. D
27. A
28. A
29. C
30. A

EQUILIBRIUM HL PAPER 1B

1. (a) Math steps assigning initial and change properly (e.g. Initial concentrations: $[H_2]$ = 0.50 M, $[I_2]$ = 0.50 M, [HI] = 0. Changes: -x, -x, +2x.) resulting in Equilibrium expressions: 0.50-x, 0.50-x, 2x. [3] (1 for init, 1 for change, 1 for eq).

(b) $K_c = (2x)^2 / (0.50-x)^2 = 54.0$. Square root both sides: $2x / (0.50-x) = 7.348$. x = 0.393 M [1]. [HI] = 2(0.393) = 0.786 mol dm^{-3} [1].

2. (a) $\Delta G^\ominus = -RT \ln K$ => $4730 = -8.31 \times 298 \times \ln K$ [1]. $\ln K = -1.91$ => $K = 0.148$ [1].

(b) The equilibrium lies to the left; the mixture is predominantly reactants [1].

(c) For an endothermic reaction, increasing T increases K [1]. Since K becomes larger (closer to 1 or >1), ln K increases, meaning $\Delta G^\ominus$ becomes less positive or negative (more spontaneous) [1].

3. (a) To get $4X \rightleftharpoons Z$, we need 2 × Reaction I ($4X \rightleftharpoons 2Y$). $K_new1 = (K_1)^2 = 16.0$ [1]. We then reverse Reaction II ($2Y \rightleftharpoons Z$). $K_new2 = 1 / K_2 = 1 / 0.50 = 2.0$ [1]. Add the two adjusted reactions to get $4X \rightleftharpoons Z$. $K_3 = K_new1 \times K_new2 = 16.0 \times 2.0 = 32$ [1].

4. (a) The volume of the container was suddenly increased (pressure was decreased) [1]. This is evidenced by the simultaneous, immediate drop in concentration of all three gases at t = 10 s (a dilution effect) [1].

(b) Upon sudden volume increase, the partial pressures drop. Q_c involves the squared product terms over cubed/linear reactant terms. A uniform decrease in concentration makes $Q_c > K_c$ mathematically [1]. According to Le Chatelier, the system shifts to the left (side with more gas moles: 4 vs 2) to restore equilibrium pressure [1], consuming NH_3 gradually [1].

EQUILIBRIUM HL PAPER 2

1. (a) $K_c = [N_2O_4] / [NO_2]^2$ [1]. $= 0.800 / (0.200)^2 = 20.0$ [1].

(b) Decreasing temperature shifts equilibrium favorably to the exothermic direction [1]. Because ΔH is negative (exothermic right), the system shifts right, increasing K_c [1]. The shift converts 2 moles of gas into 1 mole. Thus the total number of gas molecules decreases, reducing the total pressure [1].

2. (a) Initial $COCl_2 = 0.400 / 2.00 = 0.200$ M. Change in Cl_2 is +0.120 M, so change in $COCl_2$ is -0.120 M [1]. Eq $[COCl_2] = 0.200 - 0.120 = 0.080 \text{ mol dm}^{-3}$ [1].

(b) $K_c = [CO][Cl_2] / [COCl_2] = (0.120)(0.120) / (0.080)$ [1]. $K_c = 0.18$ [1].

3. (a) $K_c = [CO_2]$ [1]. Solid concentrations are constant, so they are physically omitted from heterogenous equilibrium expressions (absorbed into the K_c constant value) [1].

4. (a) For an exothermic reaction, ΔH is negative, making the term $(-\Delta H / RT)$ positive [1]. As T increases, the magnitude of 1/T decreases, so the positive term becomes smaller [1]. Thus, ln K decreases, meaning K decreases [1].

5. (a) K_c is extremely small (10^{-5}), so the equilibrium point lies far to the left [1]. The amount dissociated (x) will be vastly smaller than 0.10, so $0.10 - x \approx 0.10$ [1].

6. (a) $a < b$ [1]. (Shift to left upon compression implies left side has fewer moles).

7. (a) Macroscopically, there are no observable changes as the system is at equilibrium [1]. Microscopically, the forward and reverse reactions are continuously occurring at exactly equal rates [1].