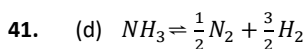


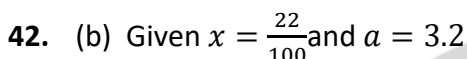
K_p & K_c Relationship and Characteristics of K



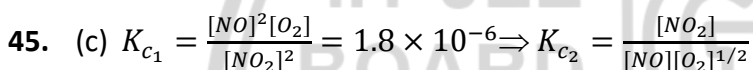
$$K_c = \frac{[N_2]^{1/2}[H_2]^{3/2}}{[NH_3]} \text{ and } \frac{1}{2}N_2 + \frac{3}{2}H_2 \rightleftharpoons NH_3$$

$$K_c = \frac{[NH_3]}{[N_2]^{1/2}[H_2]^{3/2}}$$

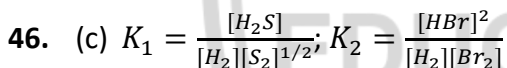
$$\text{So for dissociation} = \frac{1}{K_c}$$



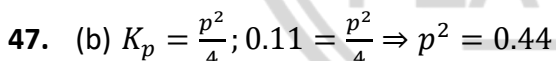
$$\therefore [HI] \text{ at equilibrium} = 3.2 \left[1 - \frac{22}{100} \right] = 2.496$$



$$K_{c1} = \frac{1}{K_{c2}^2}; 1.8 \times 10^{-6} = \frac{1}{K_{c2}^2} \Rightarrow K_{c2} = 7.5 \times 10^2$$



$$K_3 = \frac{[HBr]^2 \times [S_2]^{1/2}}{[Br_2] \times [H_2S]}; \frac{K_2}{K_1} = K_3$$



$$\text{or } p = \sqrt{0.44} = 0.66444 \approx 0.665 \text{ atm}$$



Why: Volume changes affect equilibria when the change in moles of gas, $\Delta n = (\text{moles of gaseous products}) - (\text{moles of gaseous reactants})$, is nonzero.

- (a) $\Delta n = (1 + 3) - 2 = +2 \rightarrow$ affected.
- (b) $\Delta n = 1 - (1 + 1/2) = -1/2 \rightarrow$ affected.
- (c) $\Delta n = 1 - 2 = -1 \rightarrow$ affected.

Since none have $\Delta n = 0$, all are affected by a volume change.



- 49 (a) $K_2 \approx 0.25 K_1$

Reason (word-friendly steps):

Use Arrhenius ratio: $K_2/K_1 = \exp[-E_a/R \cdot (1/T_2 - 1/T_1)]$

Plugging values: $1/T_2 - 1/T_1 = 1/280 - 1/300 = (300 - 280)/(280 \cdot 300) = 20/84,000 = 0.000238095...$

$E_a/R \approx 11570 / 1.987 \approx 5824.5$

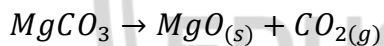
So exponent $\approx -5824.5 \times 0.000238095 \approx -1.3863$

Hence $K_2/K_1 \approx e^{-1.3863} \approx 0.25$

So K_2 is about one quarter of K_1 .

50. (a) $C_{12}H_{22}O_{11(s)} + 12O_{2(g)} \rightarrow 12CO_{2(g)} + 11H_2O$
 $\Delta n = 12 - 12 = 0$

51. (a) In this reaction gaseous molecule count



$$K_p = P_{CO_2}$$

- 52 (b) K^2

Word-friendly explanation:

When a balanced chemical equation is multiplied by a factor n , the new equilibrium constant becomes K^n .

Here, the new equation is obtained by multiplying the original equation by 2.

So,

$$K' = (K)^2 = K^2.$$

53. (a) $K_p = K_c[RT]^{\Delta n}$; $\Delta n = -1$, $K_c = 26$

$$R = 0.0812, T = 250 + 273 = 523K$$

$$K_p = 26[0.0812 \times 523]^{-1} = 0.605 \approx 0.61$$



54 (a) Unchanged

Word-friendly explanation:

The **equilibrium constant (K_p)** depends **only on temperature**, not on pressure, concentration, or catalyst.

So even if the pressure is increased tenfold, **K_p remains the same**, unless **temperature changes**.

55. (a) $2AB \rightleftharpoons A_2 + B_2$

$$K_c = \frac{[A_2][B_2]}{[AB]^2}$$

For reaction $AB \rightleftharpoons \frac{1}{2}A_2 + \frac{1}{2}B_2$

$$K_c' = \frac{[A_2]^{1/2}[B_2]^{1/2}}{[AB]}; K_c' = \sqrt{K_c} = \sqrt{49} = 7.$$

56. (a) For this reaction Δn is negative & ΔH is positive so it take forward by decrease in temperature.

57. (b) Chemical equilibrium of reversible reaction is not influenced by catalyst. It is affected by pressure, temperature & concentration of reactant.

58. (c) ΔH is positive so it will shift toward the product by increase in temperature.

59. (c) $CO(g) + Cl_2(g) \rightleftharpoons COCl_2(g)$

$$\Delta n = 1 - 2 = -1$$

$$K_p = K_c[RT]^{\Delta n}; \therefore \frac{K_p}{K_c} = [RT]^{-1} = \frac{1}{RT}$$

60. (c) ΔH is positive so reaction move forward by increase in temperature & value of $\Delta n = 3 - 2 = +1$ is positive so it forward with decrease in pressure.

