09 Chemical Equilibria

Set A **Answer key**

A From the lesson

13 MARKS

1

$K_c = \frac{[NH_3]^2}{[N_2][H_2]^3}$ $mol^{-2}\ dm^6$ [1]

$K_p = \frac{p_{NH_3}{}^2}{p_{N_2}\ p_{H_2}{}^3}$ atm^{-2} [1]

2

$K_c = \frac{[H^+]^3}{[Fe^{3+}]}$ $mol^2\ dm^{-6}$: leave out the solid and the solvent, water. [1]

3

Since 1.56 mol of HI is formed, the amounts of H_2 and I_2 each fall by 0.78 mol, leaving 0.22 mol of each. [1]

$K_c = \frac{(1.56)^2}{(0.22)(0.22)}$ = **50.3** (no units) [1]

4 **(a)**

The position of equilibrium shifts to the right, as there are fewer moles of gas on the right (2 mol, against 4 mol on the left). [1]

(b)

The position of equilibrium does not shift, as there are equal numbers of moles of gas on both sides (2 mol each). [1]

(c)

The position of equilibrium shifts to the left, as there are fewer moles of gas on the left (1 mol, against 2 mol on the right). [1]

5

Rate of forward reaction equals rate of backward reaction, so there is no net change in concentrations, in a closed system. [1]

6

$K_c = \frac{[SO_3]^2}{[SO_2]^2[O_2]}$, $mol^{-1}\ dm^3$ [1]

7

It increases the rates of the forward and backward reactions equally, so K and the POE are unchanged. [1]

8

A lower T gives a higher yield, but the rate is too slow [1]

450 °C gives a reasonable yield in a reasonable time. [1]

B Practice

21 MARKS

9 **A** 2 only [1]

Temperature is the only factor that will change an equilibrium constant, K_c or K_p.

10 **A** removal of H_2O [1]

To favour the oxidation is to favour the forward reaction (the position of equilibrium shifts to the right). Removal of H_2O will shift the position of equilibrium to the right.

B Adding a catalyst will not affect the position of equilibrium.

C This is an exothermic reaction, hence increasing the temperature will shift the position of equilibrium to the left, favouring the endothermic backward reaction to absorb the additional heat.

D Reducing the volume will increase the pressure, and to decrease the pressure, the position of equilibrium will shift to the left to produce fewer gaseous molecules.

11 **(a)**

$$K_c' = \frac{[X_2Y]^4}{[X_2]^4[Y_2]^2} = (K_c)^2 = (5.00)^2$$

$= \mathbf{25.0\ mol^{-2}\ dm^6}$ [1]

(b)

$$K_c'' = \frac{[X_2][Y_2]^{½}}{[X_2Y]} = \left(\frac{1}{K_c}\right)^{½} = \left(\frac{1}{5.00}\right)^{½}$$

$= \mathbf{0.447\ mol^{½}\ dm^{-3/2}}$ [1]

12

Amount of NaOH reacted with $CH_3COOH = 0.200 \times \frac{289}{1000} = 0.0578$ mol = amount of CH_3COOH present [1]

	CH_3COOH	C_2H_5OH	$CH_3COOC_2H_5$	H_2O
initial amount / mol	1.00	5.00	0	0
change / mol	−0.9422	−0.9422	+0.9422	+0.9422
equilibrium amount / mol	0.0578	4.0578	0.9422	0.9422

Let the total volume be V dm^3. $K_c = \frac{[H_2O][CH_3COOC_2H_5]}{[CH_3COOH][C_2H_5OH]} = \frac{(0.9422/V)(0.9422/V)}{(0.0578/V)(4.0578/V)}$ [1]

$=$ **3.79** (no units) [1]

13 **(a)**

$K_p = \frac{(p_{CO})(p_{H_2})}{p_{H_2O}}$, unit of K_p: Pa *or* atm [1]

(b)

When pressure is increased, the position of equilibrium will shift to decrease pressure by decreasing the total number of gaseous molecules present. Since there are 2 molecules of gas on the right-hand side of the equation and one on the left, the backward reaction is favoured. The position of equilibrium shifts to the left. Hence, at equilibrium, there will be more H_2O, and less H_2 and CO. **[1]**

When temperature is increased, the position of equilibrium will shift to decrease temperature. Since the forward reaction is endothermic, the forward reaction is favoured. The position of equilibrium shifts to the right. Hence, at equilibrium, there will be more H_2 and CO, and less H_2O. **[1]**

(c)

Since the amount of H_2 produced = the amount of CO produced, the partial pressures of the two gases must be the same. Partial pressure of CO = **183 kPa** **[1]**

$K_p = \frac{(183\ \text{kPa})(183\ \text{kPa})}{90\ \text{kPa}}$ = **372 kPa** **[1]**

Let the new partial pressure of H_2 be x kPa. The partial pressure of CO is also x kPa. $K_p = \frac{(x\ \text{kPa})(x\ \text{kPa})}{150\ \text{kPa}} = 372$ kPa, so $x = \sqrt{(372 \times 150)}$ = **236 kPa** (the negative root is disregarded, as partial pressures cannot be negative) **[1]**

14 **(a)**

	Zn(s)	$H_2O(g)$	ZnO(s)	$H_2(g)$
initial amount / mol	2.60	2.60	0	0
change in amount / mol	−2.58	−2.58	+2.58	+2.58
equilibrium amount / mol	0.02	0.02	2.58	2.58

Partial pressure of H_2 at equilibrium = $(\frac{2.58}{0.02 + 2.58}) \times 10$ = **9.92 atm** **[1]**

(b)

$K_p = \frac{p_{H_2}}{p_{H_2O}}$ **[1]**

$= \frac{(2.58/2.60) \times 10}{(0.02/2.60) \times 10}$ = **129** **[1]**

(c)

At lower temperatures, the position of equilibrium shifts right to favour the forward exothermic reaction that releases heat, increasing the yield of zinc oxide. **[1]**

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$K_p = \frac{(p_{CH_4})(p_{H_2O})}{(p_{CO})(p_{H_2})^3}$ **[1]**

Let the initial amounts of CO and H_2 be 1 and 3 mol respectively.

	CO	$3H_2$	CH_4	H_2O
initial amount / mol	1	3	0	0
change / mol	−0.40	−3(0.40)	+0.40	+0.40
equilibrium amount / mol	0.60	1.80	0.40	0.40

Total amount at equilibrium = 0.6 + 1.8 + 0.4 + 0.4 = 3.2 mol **[1]**

$p_{CO} = \frac{0.6}{3.2} \times 32 = 6$ atm, $p_{H_2} = \frac{1.8}{3.2} \times 32 = 18$ atm, $p_{CH_4} = p_{H_2O} = \frac{0.4}{3.2} \times 32 = 4$ atm **[1]**

$K_p = \frac{(4)(4)}{(6)(18)^3} =$ **4.57×10^{-4} atm^{-2}** **[1]**