Chapter 8: Equilibrium

Homework marking scheme

1 a i $\Delta H_{reaction} = \Delta H_f(\text{products}) - \Delta H_f(\text{reactants})$ [1]
$= 2(+33.2) - 9.2$ [1]
$= +57.2 \text{ kJ mol}^{-1}$ [1]

ii The equilibrium will shift to the right-hand side [1]
because the forward reaction is endothermic. [1]
Heating favours the reaction that cools down (Le Chatelier's principle) [1]

iii Equilibrium shifts to the right-hand side and the pressure of the NO_2 will increase. [1]
Therefore, K_p will increase. [1]

b Any **two** from:
The system is a closed system.
The rate of forward reaction = rate of backward reaction.
The macroscopic properties, pressure, temperature, etc. are constant. [2]

c i $K_p = \dfrac{P^2(NO_2)}{P(N_2O_4)}$ [1]

units = Pa, because $K_p = \dfrac{(\text{pressure})^2}{\text{pressure}}$ [1]

ii $N_2O_4(g) \rightleftharpoons 2NO_2(g)$

	$N_2O_4(g)$	$2NO_2(g)$	Explanation
At start	1 mol	0 mol	
At equilibrium	1 – 0.8 mol = 0.2 mol	1.6 mol	for every mol of N_2O_4 converted, 2 mol of NO_2 are formed total quantity of gases = 1.8 mol

[2]

partial pressure of $N_2O_4 = 1 \times 10^5 \times \dfrac{0.2}{1.8} = 0.111 \times 10^5$ Pa

partial pressure of $NO_2 = (1 - 0.111) \times 10^5 \text{ Pa} = 0.889 \times 10^5$ Pa [1]

$K_p = \dfrac{(0.889 \times 10^5)^2}{0.111 \times 10^5}$ [1]

$= 7.12 \times 10^5$ Pa [1]

iii In this question K_p remains constant because the temperature remains constant, only P changes

	$N_2O_4(g)$	$2NO_2(g)$	Explanation
At start	1 mol	0 mol	
At equilibrium	1 – 0.2 mol = 0.8 mol	0.4 mol	For every mol of N_2O_4 converted, 2 mol of NO_2 are formed total quantity of gases = 1.2 mol

[2]

partial pressure of $N_2O_4 = \dfrac{0.8}{1.2} \times P = 0.667P$

partial pressure of $NO_2 = \dfrac{0.4}{1.2} \times P = 0.333P$ [1]

$K_p = \dfrac{(0.333P)^2}{0.667P} = 7.12 \times 10^5$ Pa [1] new pressure = 42.6×10^5 Pa [1]

iv The new pressure is higher than before, which would favour the side of the equilibrium with fewer gas molecules. [1]

Therefore, there is a lower conversion of N_2O_4 into NO_2. [1]

d i The oxidation state of nitrogen in $NO_2 = +4$ [1]

in $NaNO_2$ it is +3, therefore reduced [1]

in $NaNO_3$ it is +5, therefore oxidised; hence, a redox reaction. [1]

ii number of mol of NaOH = number of mol of $NO_2 = \frac{480}{24000} = 0.02$ mol [1]

$V = \frac{n}{C} = \frac{0.02}{0.25}$ [1]

$= 0.0800\ \text{dm}^3$ or $80\ \text{cm}^3$ [1]

2 a $K_p = \frac{P^2(NH_3)}{P(N_2) \times P^3(H_2)}$ [2]

Unit is Pa^{-2} because $K_p = \frac{(\text{pressure})^2}{(\text{pressure})^4}$ [1]

b

	N_2(g)	$3H_2$(g)	$2NH_3$(g)	
Number of mol	1 – 0.15 = 0.85	3 – 0.45 = 2.55	0.30	total number of mol = 3.7
Partial pressures	$\frac{0.85}{3.7} \times 2 \times 10^7$ = 4.59×10^6 Pa	$\frac{2.55}{3.7} \times 2 \times 10^7$ = 1.38×10^7 Pa	$\frac{0.30}{3.7} \times 2 \times 10^7$ = 1.62×10^6 Pa	

[3]

Therefore, $K_p = \frac{(1.62 \times 10^6)^2}{(4.59 \times 10^6) \times (1.38 \times 10^7)^3} = 2.18 \times 10^{-16}\ \text{Pa}^{-2}$ [2]

c The extent to which the equilibrium moves to the right-hand side (the products). [1]

d A decrease in K_p means a lower yield of products. [1]

Increasing the temperature favours the backward reaction. [1]

Therefore, the backward reaction is endothermic. [1]

The forward reaction is exothermic. [1]

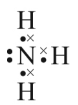
e i The yield of ammonia would decrease [1]

because the decrease in pressure would favour the side of the reaction with more molecules to counteract the decrease in pressure (Le Chatelier's principle). [1]

ii K_p does not change. [1]

Only temperature affects the value of K_p. [1]

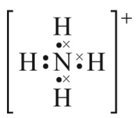
f i



three dot-cross bonds [1]

the lone pair on the nitrogen [1]

ii

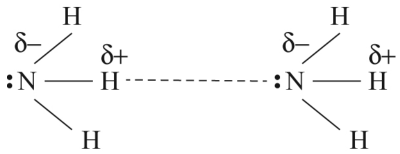


three dot-cross bonds [1]

dative covalent bond with fourth hydrogen [1]

g The lone pair on the nitrogen in ammonia repels the covalent bonds more than the dative covalent bond in the ammonium ion. [1]
Therefore, the bond angle in ammonia is less than in the ammonium ion. [1]

h



correct dipoles [1]
lone pair electrons on the nitrogen with the hydrogen bond [1]
hydrogen bond as a dashed line. [1]

3 **a** $+90\ \text{kJ mol}^{-1}$ [1]
2 mol of NO are formed from its constituent elements in their standard states. Therefore, divide by 2. [1]

b **i** $K_c = \dfrac{[NO]^2}{[N_2][O_2]}$ [1]

No units [1]

ii $[NO]^2 = [N_2]\ [O_2] \times K_c$

$[NO] = \sqrt{(3.59\times10^{-2})\times(8.42\times10^{-3})\times(4\times10^{-31})} = 1 \times 10^{-17}\ \text{mol dm}^{-3}$

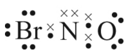
rearranging formula to give [NO] [1]
correct answer [1]
units. [1]

iii $[NO] = 1 \times 10^{-17}\ \text{mol dm}^{-3} = 1 \times 10^{-17} \times 10^{3}\ \text{mol m}^{-3}$
$= 6.023 \times 10^{23} \times 1 \times 1 \times 10^{-17} \times 10^{3}$ particles $= 6.023 \times 10^{9}$ particles per m^3
converting the number of moles per dm^3 to per m^3 [1]
finding the number of particles using Avogadro's constant. [1]

c **i** $2NO + Br_2 \rightarrow 2NOBr$

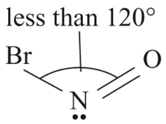
ii nitrogen: its oxidation state increases from +2 to +3, therefore oxidised [1]
bromine: its oxidation decreases from 0 to –1, therefore reduced; combination means a redox reaction. [1]

iii



double bond between N and O (two pairs of electrons) [1]
single N–Br bond (one pair of electrons) [1]
all three atoms having outer octet. [1]

iv



[1]
The shape is non-linear [1]
around the central nitrogen the electron count is: N = 5e; O = 2e; Br = 1e [1]
less 2e for the double bond gives a total of 6e, in three pairs. [1]
The bond angle is less than 120° because of the lone pair on the central nitrogen. [1]