

Marking scheme for Core Worksheet – Chapter 7

- 1** macroscopic properties constant / all concentrations remain constant [1]
rate of forward reaction = rate of reverse reaction [1]
- 2** **a** position of equilibrium shifts to left [1]
more moles of gas on right [1]
b position of equilibrium shifts to right [1]
more moles of gas on left [1]
c position of equilibrium shifts to left [1]
more moles of gas on right [1]
- 3** **a** position of equilibrium shifts to right [1]
position of equilibrium shifts in exothermic direction (to give out heat to minimise the effect of the change) [1]
b position of equilibrium shifts to right [1]
position of equilibrium shifts in exothermic direction (to give out heat to minimise the effect of the change) [1]
c position of equilibrium shifts to left [1]
position of equilibrium shifts in exothermic direction (to give out heat to minimise the effect of the change) [1]
- 4** **a** position of equilibrium shifts to right, therefore more orange [1]
to use up the added $\text{H}^+(\text{aq})$ [1]
b position of equilibrium shifts to left, therefore more yellow [1]
 OH^- reacts with H^+ [1]
position of equilibrium shifts to restore H^+ [1]
- 5** less soluble [1]
position of equilibrium shifts to left as higher concentration of OH^- present [1]
- 6** percentage dissociation will be lower [1]
extra H^+ present from sulfuric acid [1]
position of equilibrium shifts to left to use up the extra H^+ [1]

- 7 a $K_c = \frac{[\text{BX}(\text{g})]^2[\text{AB}(\text{g})]}{[\text{X}_2(\text{g})][\text{AB}_3(\text{g})]}$ [1]
- b $K_c = \frac{[\text{CH}_4(\text{g})][\text{H}_2\text{O}(\text{g})]}{[\text{CO}(\text{g})][\text{H}_2(\text{g})]^3}$ [1]
- c $K_c = \frac{[\text{CO}(\text{g})][\text{H}_2(\text{g})]^3}{[\text{CH}_4(\text{g})][\text{H}_2\text{O}(\text{g})]}$ [1]
- d $K_c = \frac{[\text{CO}(\text{g})][\text{H}_2(\text{g})]^2}{[\text{CH}_3\text{OH}(\text{g})]}$ [1]
- e $K_c = \frac{[\text{SO}_2(\text{g})]^2[\text{O}_2(\text{g})]}{[\text{SO}_3(\text{g})]^2}$ [1]
- 8 a K_c increases [1]
position of equilibrium shifts to right [1]
in the exothermic direction [1]
- b no change [1]
only temperature affects the value of the equilibrium constant [1]
- c no change [1]
only temperature affects the value of the equilibrium constant [1]
- 9 a $\text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightarrow 2\text{NH}_3(\text{g})$ [1]
- b $K_c = \frac{[\text{NH}_3(\text{g})]^2}{[\text{N}_2(\text{g})][\text{H}_2(\text{g})]^3}$ [1]
- c pressure about 200 atm [1]
temperature about 500 °C [1]
iron catalyst [1]
high pressure – gives greater yield of NH_3 [1]
fewer moles of gas on right-hand side [1]
so high pressure also speeds up reaction – particles collide more frequently [1]
even higher pressure too expensive [1]
compromise temperature between yield and rate [1]
equilibrium yield higher at lower temperature as reaction is exothermic [1]
but lower temperature gives much slower rate [1]
catalyst speeds up rate of reaction [1]
catalyst allows a lower temperature to be used, which saves money [1]



- 10 a** higher temperature gives a much higher rate of reaction [1]
the reaction is exothermic, therefore a lower temperature should favour the forward reaction and produce a better equilibrium yield of SO₃ [1]
- b** although a higher pressure would give a better yield [1]
the yield is already high enough at atmospheric pressure [1]