

Pages 17–19 Exam practice questions

- 1 a) i) $K_c = \frac{[\text{Sn}^{4+}(\text{aq})][\text{Fe}^{2+}(\text{aq})]^2}{[\text{Sn}^{2+}(\text{aq})][\text{Fe}^{3+}(\text{aq})]^2}$ [2]
- ii) K_c has no units [1]; the concentration terms cancel – $(\text{mol dm}^{-3})^3$ appears on the top and the bottom of the expression for K_c . [1]
- b) i) The reciprocal of the value given in the question [1]: $K_c = 1 \times 10^{-10}$ [1]
- ii) The square root of the value given in the question [1]: $K_c = 1 \times 10^5$ [1]
- 2 a) $K_c = \frac{(0.070 \text{ mol dm}^{-3})^2}{(0.010 \text{ mol dm}^{-3}) \times (0.010 \text{ mol dm}^{-3})}$ [1]
- = 49 [1]; no units [1]
- b) i) The quantity of iodine that has reacted = $0.0030 \text{ mol dm}^{-3}$
This reacts with $0.0030 \text{ mol dm}^{-3}$ hydrogen and produces $0.0060 \text{ mol dm}^{-3}$ HI.
So the new concentrations are:
 $[\text{H}_2(\text{g})] = 0.017 \text{ mol dm}^{-3}$ [1] and $[\text{HI}(\text{g})] = 0.076 \text{ mol dm}^{-3}$ [1].
- ii) Substituting in the expression for K_c gives:
- $$\frac{(0.076 \text{ mol dm}^{-3})^2}{(0.017 \text{ mol dm}^{-3}) \times (0.0070 \text{ mol dm}^{-3})} = 48.5$$
- [1]
- This is very close to the value for K_c calculated in (a). The system is again at equilibrium. [1]
- c) Doubling the hydrogen concentration causes a shift in the position of equilibrium to the right giving more HI and less iodine [1]; some, but not all, of the added hydrogen reacts. [1]
- 3 a) At equilibrium the forward and backward reactions happen at the same rate so that overall there is no change. [1] Nitrogen and hydrogen react to form ammonia as fast as ammonia decomposes back to nitrogen and hydrogen. [1]
- b) i) Mole fractions: NH_3 , 0.137 [1]; H_2 , 0.655 [1]; N_2 , 0.209 [1]
- ii) Partial pressures: NH_3 , 1.37 atm; H_2 , 6.55 atm; N_2 , 2.09 atm [2]
- iii) $K_p = \frac{(\rho_{\text{NH}_3})^2}{\rho_{\text{N}_2} \times (\rho_{\text{H}_2})^3}$ [1]
- = 3.20×10^{-3} [1]; atm^{-2} [1]
- 4 a) Chemical equilibrium is dynamic [1]; on a molecular scale there is constant change with opposing reactions cancelling each other out. [1] However, the overall macroscopic effect is that nothing appears to be changing when the system is at equilibrium. [1]
- b) K_c does not vary with concentration. [1] However, it is true that adding more of a reactant to a system at equilibrium means that the mixture is no longer at equilibrium [1]; some of the added chemical reacts and more products form until the system is again at equilibrium. [1]
- c) Adding a catalyst does not affect the value of K_c . [1] At a constant temperature, the yield of product at equilibrium is not changed by adding a catalyst [1]; however, many processes are

so slow that they never reach equilibrium in the absence of a catalyst, so in that sense an added catalyst increases the yield by enabling the reaction mixture to reach equilibrium. [1]

- d) This is true if the reaction to make products is endothermic. [1] If the reaction is exothermic, then the equilibrium yield of product falls as the temperature rises. [1] However, as with catalysts, raising the temperature can increase the yield of a product by allowing a reaction mixture to reach equilibrium that reacts too slowly at room temperature. [1]
- e) A reaction is feasible in the thermodynamic sense if the total entropy change for the process is positive. [1] Adding a catalyst speeds up the rate of reaction but cannot change the feasibility of a reaction at a particular temperature. [1] Only if the reaction is not going because it is kinetically inert can a catalyst enable a reaction to happen at a particular temperature. [1]

5 a) Concentration in aqueous layer = $\frac{17}{10} \times 0.50 \text{ mol dm}^{-3}$ [1]

$$= 0.85 \text{ mol dm}^{-3} \text{ [1]}$$

b) Concentration in organic solvent = $\frac{6.0}{10} \times 0.01 \text{ mol dm}^{-3}$ [1]

$$= 0.006 \text{ mol dm}^{-3} \text{ [1]}$$

c) $K_c = \frac{0.85 \text{ mol dm}^{-3}}{0.006 \text{ mol dm}^{-3}} = 140$ [1]; no units [1]

6 a) $K_c = \frac{[\text{PCl}_3(\text{g})][\text{Cl}_2(\text{g})]}{[\text{PCl}_5(\text{g})]}$ [1]

b) $K_c = \frac{1}{8 \times 10^{-3} \text{ mol dm}^{-3}} = 125$ [1] $\text{dm}^3 \text{ mol}^{-1}$ [1]

c) $[\text{PCl}_3(\text{g})] = [\text{Cl}_2(\text{g})] = x$ [1]

$$\frac{x^2}{5 \times 10^{-2} \text{ mol dm}^{-3}} = 8 \times 10^{-3} \text{ mol dm}^{-3} \text{ [1]}$$

$$x = 2 \times 10^{-2} \text{ mol dm}^{-3} \text{ [1]}$$

- d) i) Some, but not all, of the added PCl_5 decomposes [1] and increases the concentrations of PCl_3 and chlorine. [1]
- ii) Increasing the pressure causes the equilibrium to shift in the direction that gives fewer moles of gas [1]; so the concentration of PCl_5 increases while the concentrations of PCl_3 and chlorine decrease. [1]
- iii) Raising the temperature causes the equilibrium to shift in the direction that is endothermic [1]; so the concentration of PCl_5 decreases while the concentrations of PCl_3 and chlorine increase. [1]
- e) i) Changing concentrations does not affect the value of K_c . [1]
- ii) Changing the pressure does not affect the value of K_c [1]
- iii) K_c increases [1] For an endothermic reaction the value of K_c gets larger as the temperature rises [1]

- 7 a) $K_p = \frac{p_{\text{CO}} \times (p_{\text{H}_2})^3}{p_{\text{CH}_4} \times p_{\text{H}_2\text{O}}}$ [2]
- b) i) K_p does not change [1]
 ii) K_p increases [1]
 iii) K_p does not change [1]
- c) i) Concentration of products decreases and the concentration of reactants increases. [1]
 ii) Concentration of products increases and the concentration of reactants decreases. [1]
 iii) Composition of the equilibrium mixture does not change. [1]
- 8 a) $4\text{NH}_3(\text{g}) + 5\text{O}_2(\text{g}) \rightleftharpoons 4\text{NO}(\text{g}) + 6\text{H}_2\text{O}(\text{g})$ [2]
- b) In theory the yield of NO at equilibrium is high if the temperature is low [1] and the pressure low [1]; adding excess oxygen also favours the conversion of ammonia to NO. [1]
- c) There is an excess of oxygen (even though only a fifth of the air is oxygen). [1] The temperature has to be high because reaction only happens at a practicable rate if the catalyst is hot. [1] Raising the pressure compresses the gases and this increases the mass of material flowing through the reactor and hence the mass of product produced in a given time from a given piece of plant. [1]
 (The chosen pressure also depends on the price of platinum. Working at higher pressures increases the wear on the catalyst and shortens the life of the catalyst gauzes.)
- d) The gases flow through the solid catalyst gauzes and so the catalyst is used continuously and does not have to be recovered from the reaction mixture. [1]
- e) The gases flow through the catalyst gauzes so the mixture of product and unchanged reactants is only in contact with the catalyst for a short time [1]; as soon as the gases flow away from the catalyst reaction stops – the mixture has time to approach equilibrium but not reach equilibrium. [1]
- f) The hot gases flow through a heat exchanger where they turn water to hot steam [1]; that can be used in other processes or to generate electricity. [1]

9 a)

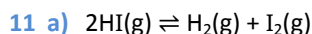
	$\text{COCl}_2(\text{g}) \rightleftharpoons \text{CO}(\text{g}) + \text{Cl}_2(\text{g})$		
Initial amount/mol	1	0	0
Amount at equilibrium/mol	$(1 - \alpha)$	α	α
Equilibrium partial pressure/Pa	$\frac{(1 - \alpha)P}{(1 + \alpha)}$	$\frac{\alpha P}{(1 + \alpha)}$	$\frac{\alpha P}{(1 + \alpha)}$

Terms in second row of table [1]; total amounts in moles [1]; mole fractions [1]; partial pressures. [1]

$$K_p = \frac{\alpha^2 P}{(1 + \alpha)(1 - \alpha)} \quad (\text{working [1], answer [1]})$$

- b) i) $\alpha = 1.4 \times 10^{-3}$ (working [1], answer [1])
 ii) $\alpha = 1.0 \times 10^{-3}$ (working [1], answer [1])

- 10 a) i) Equal volumes of gases contain equal amounts, in moles, of gas. So the mole ratio is 1 : 1. [1] According to the equation 1 mol SO₂ reacts with 0.5 mol O₂, so the oxygen is in excess. [1]
- ii) According to the equilibrium law (and Le Chatelier's principle), adding an excess of a reactant tends to increase the conversion of other reactants to products. [1] Excess oxygen favours the formation of sulfur trioxide. [1]
- iii) The oxygen comes from the air and is mixed with nitrogen. [1] Adding more air would dilute the sulfur dioxide so much that too little of the gas would be flowing into the plant to produce product at a reasonable rate. [1]
- b) 700 K is hot enough for the catalyst to be effective and for the reaction to be fast enough. [1] The temperature is not higher because the reaction is exothermic and so the value of K_p falls as the temperature rises. [1] The equilibrium yield of product falls as the temperature rises. [1]
- c) Raising the pressure would increase the yield of product at equilibrium because 3 mol reactants combine to give 2 mol product. [1] Increasing the pressure would also increase the mass of chemicals flowing through the plant. However a high-pressure plant costs more to build and raises the running costs because of the energy needed to run the pumps. [1] Since the conversion to SO₃ is close to 100% it is generally not worth operating at a higher pressure. [1]
- d) The reaction is exothermic so at each stage the gas mixture tends to heat up. [1] The gas mixture is cooled between catalyst beds to ensure that the equilibrium yield of product is as high as possible. [1] (The energy is used to raise steam, which can be used in other processes or used to generate electricity.)
- e) Removing the product from the gas stream when the sulfur dioxide concentration is low helps to shift the equilibrium to the right [1] to ensure that almost all the sulfur dioxide is converted to sulfur trioxide. [1]
- f) Sulfur dioxide is an acidic gas which causes acid rain. It is not acceptable to release any traces of sulfur dioxide into the air. [1] There is a cost to removing the last traces of sulfur dioxide from the remaining gases (which are mainly nitrogen and excess oxygen). [1]



Initial amounts: $\text{HI}(\text{g}) = 0 \text{ mol}$, $\text{H}_2(\text{g}) = 2 \text{ mol}$ and $[\text{I}_2(\text{g})] = 1 \text{ mol}$

Let the equilibrium amount of HI = $2x$

Then the equilibrium concentrations:

$$[\text{HI}(\text{g})] = \frac{2x}{1 \text{ mol dm}^{-3}}, \quad [\text{H}_2(\text{g})] = \frac{(2-x)}{1 \text{ mol dm}^{-3}} \quad \text{and} \quad [\text{I}_2(\text{g})] = \frac{(1-x)}{1 \text{ mol dm}^{-3}} \quad [1]$$

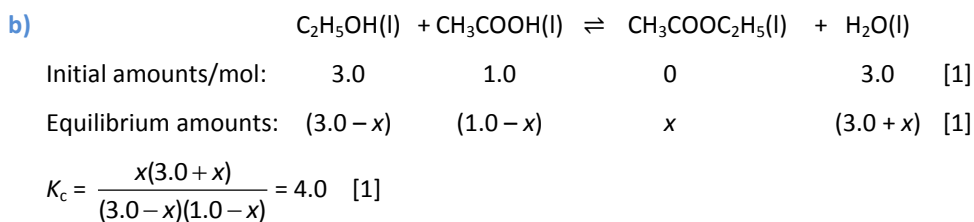
$$K_c = \frac{(2-x)(1-x)}{4x^2} = 0.020 \quad [1]$$

$$\text{Hence: } 0.92x^2 - 3x + 2 = 0.020 \quad [1]$$

Solving shows that $x = 0.935$ [1]

Hence the equilibrium concentrations are:

$$[\text{I}_2(\text{g})] = 0.065 \text{ mol}, [\text{H}_2(\text{g})] = 1.06 \text{ mol and } [\text{HI}(\text{g})] = 1.87 \text{ mol} \quad [2]$$

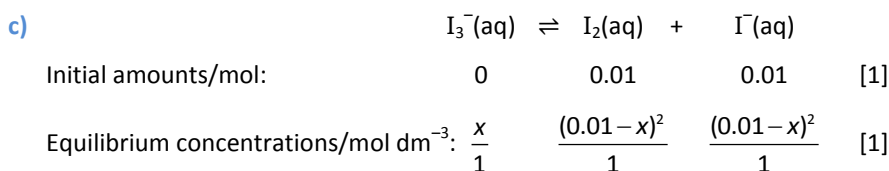


because the volume terms cancel.

$$\text{Hence: } 3x^2 - 19x + 12 = 0 \quad [1]$$

$$x = 0.71 \quad [1]$$

Amount of ethyl ethanoate at equilibrium = 0.71 mol [1]



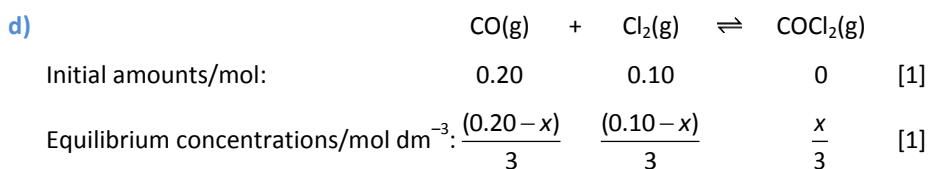
$$K_c = \frac{(0.01 - x)^2}{x} = 1.5 \times 10^{-3} \quad [1]$$

$$x^2 - 0.0215x + 0.0001 = 0 \quad [1]$$

$$x = 0.0068 \quad [1]$$

The equilibrium concentrations are:

$$[\text{I}_3^-(\text{aq})] = 0.0068 \text{ mol dm}^{-3} \quad \text{and} \quad [\text{I}^-(\text{aq})] = 0.0032 \text{ mol dm}^{-3} \quad [1]$$

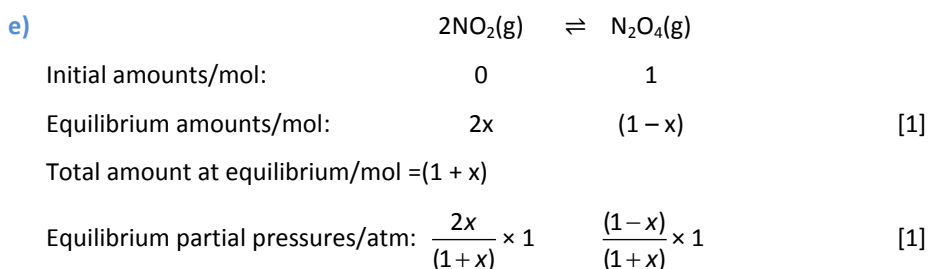


$$K_c = \frac{3x}{(0.20 - x)(0.10 - x)} = 0.41 \quad [1]$$

$$0.41x^2 - 3.123x + 0.0082 = 0 \quad [1]$$

$$x = 0.0262 \quad [1]$$

The equilibrium concentration is $[\text{COCl}_2(\text{g})] = 0.026 \text{ mol dm}^{-3}$ [1]



$$K_p = \frac{(1 - x)(1 + x)^2}{(1 + x)4x^2} = 7.1 \quad [1]$$

$$29.4x^2 - 1 = 0$$

$$x = 0.184 \quad [1]$$

$$\text{So partial pressure of N}_2\text{O}_4 = \frac{1-0.184}{1+0.184} \times 1 = 0.69 \text{ atm [1]}$$

$$\text{and partial pressure of NO}_2 = \frac{2 \times 0.184}{1+0.184} \times 1 = 0.31 \text{ atm [1]}$$