# Topic Exploration Pack

**Equilibrium and pressure – *K*p**

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## Instructions and answers for teachers

These instructions cover the student activity section which can be found on [page 21](#_Learner_activity). This Topic Exploration Pack supports OCR AS and A Level Chemistry A.

**When distributing the activity section to the students either as a printed copy or as a Word file you will need to remove the teacher instructions section.**

### Learning outcomes

5.1.2(a) use of the terms *mole fraction* and *partial pressure*

5.1.2(b) calculation of quantities present at equilibrium, given appropriate data

5.1.2(d) expressions for *K*c and *K*p for homogeneous and heterogeneous equilibria (see also 3.2.3f)

5.1.2(e) calculations of *K*c and *K*p, or related quantities, including determination of units (see also 3.2.3f)

5.1.2(f)(i) the qualitative effect on equilibrium constants of changing temperature for exothermic and endothermic reactions

(ii) the constancy of equilibrium constants with changes in concentration, pressure or in the presence of a catalyst

5.1.2(g) explanation of how an equilibrium constant controls the position of equilibrium on changing concentration, pressure and temperature

### Introduction

*K*pis the equilibrium constant in terms of partial pressures. It allows an equilibrium constant to be determined for a reaction based on the proportions of the components in a gaseous system and the overall pressure, without having to calculate the concentrations of the components – particularly useful where the volume of the container is not known or difficult to measure.

It may be tempting to view *K*p as ‘simply’ an extension of *K*c and to focus practice on calculations. However, it is important to take note that candidates can also be asked to explain the effects of changing certain factors on the position of equilibrium and on the value of the equilibrium constant. Explanation of the position of equilibrium must be in terms of the equilibrium constant; explanations in terms of Le Chatelier’s principle are not sufficient. Developing such explanations requires a good grasp of the situation, and there are particular aspects that apply to *K*p. In particular, in responding to questions about changes in pressure or volume, learners have to combine thinking about the equilibrium and the chemical reaction taking place with thinking about the physical behaviour of gases (ideal gas / Boyle’s law).

The OCR presentation ‘*[K](mailto:http://www.ocr.org.uk/Images/250387-kp-presentation.ppthttp://www.ocr.org.uk/Images/250387-kp-presentation.ppt)*[p](mailto:http://www.ocr.org.uk/Images/250387-kp-presentation.ppthttp://www.ocr.org.uk/Images/250387-kp-presentation.ppt)’ provides an overview of the key concepts in this topic. The slides could be used as a prompt when introducing learners to calculations.

**Prior knowledge**

It is likely that learners will have initially been introduced to *K*c, possibly in year 12. It makes sense to first expand *K*c into the content of Module 5, covering calculations with initial concentrations, heterogeneous equilibria, the change of the equilibrium constant with temperature and how the equilibrium constant controls the position of equilibrium. From there, you can extend the concepts into *K*p – perhaps starting with the concept of partial pressure and demonstrating the equivalence of calculations for *K*c and *K*p.

For any work on equilibrium, learners require a secure understanding of the concept of concentration. For *K*p specifically, it is advisable to review the ideal gas equation to ensure that learners understand the relationship between amount of substance (moles), pressure and volume. Using the ideal gas equation to show how pressure (*p*) is proportional to concentration (*n*/*V*­) may help learners appreciate how *K*c and *K*p are related, and thus smooth the transition from *K*c to *K*p.

The activities and worksheets in this pack are written to cover aspects of *K*p, but much of the guidance can be applied to *K*c as well.

**Common misconceptions**

* Look out for learners taking an algorithmic approach to calculations and problems without proper understanding of the concepts. Good conceptual understanding is required in order to tackle problems that do not follow the ‘usual’ pattern. It is worth asking learners to explain their thinking, and to intersperse calculations with questions requiring responses in qualitative terms rather than calculations, so that you can address any misconceptions that arise.
* Some learners may find it difficult to move beyond explanations in terms of Le Chatelier’s principle. Carefully model explanations relating to the position of equilibrium in terms of the equilibrium constant. It may help to work through numerical examples to illustrate the general qualitative points.
* Watch out for imprecise language in describing the effect of a change in conditions. For example, if an equilibrium shift following an increase in volume results in an increased *amount* of a component of a system, it is not necessarily correct that the *concentration* of that component increases.
* When explaining effects of changes in pressure or volume, learners may focus only physical effects (in terms of the ideal gas law), or put ‘Le Chatelier’ thinking in the wrong place. For example, ‘if the volume is increased, the system shifts to decrease the volume, so moves to the side with fewer moles of gas’. Here, the volume is interpreted as an attribute of the gas, rather than the container. There are many different types of errors that can be made, highlighting the importance of keeping an eye on learners’ thought processes and carefully modelling correct explanations.
* It is worth taking note of reports (e.g. Quílez, 2004[[1]](#footnote-1); Cheung, 2009[[2]](#footnote-2)) on misconceptions that can emerge from use of Le Chatelier’s principle.

### Suggested activities

A range of question sheets is provided in this pack. The order roughly follows the order in which the concepts are covered in the specification, however you may wish to use a different teaching order. For example, introducing the notion of *K*p based on partial pressures first, then working back into the concept of mole fractions.

While the activities are provided as question sheets, the content can be used flexibly rather than having learners work through the questions individually. Some ideas are provided below.

### Activity 1: Mole fractions and partial pressures

**Answers**

1. An equilibrium mixture contains 0.50 mol I2, 0.75 mol H2 and 4.5 mol HI. Calculate the mole fraction of each gas.

|  |
| --- |
| *N* = 0.50 + 0.75 + 4.5 = 5.75 mol  *x*(I2) = 0.50 / 5.75 = 0.0870  *x*(H2) = 0.75 / 5.75 = 0.130  *x*(HI) = 4.5 / 5.75 = 0.783 |

2. An equilibrium mixture contains two gases, A and B. The mole fraction of gas A is 0.46. Calculate the mole fraction of gas B.

|  |
| --- |
| *x*(B) = 1 – *x*(A) = 1 – 0.46 = 0.54 |

3. Explain why the mole fraction has no units.

|  |
| --- |
| To calculate the mole fraction, an amount in mol is divided by an amount in mol, so the units cancel out. |

**Partial pressures**

4. Look back at the equilibrium mixture described in question 1. The total pressure of this mixture is 115 kPa. Calculate the partial pressure of each gas in the mixture.

|  |
| --- |
| *p*(I2) = 0.0870 × 115 = 10.0 kPa  *p*(H2) = 0.130 × 115 = 15.0 kPa  *p*(HI) = 0.783 × 115 = 90.0 kPa |

5. An equilibrium mixture contains two gases, C and D. The mole fraction of C is 0.870. The partial pressure of C is 560 Pa. Calculate

(a) the total pressure.

|  |
| --- |
| *P* = *p*(C) / *x*(C) = 560 / 0.870 = 644 Pa |

(b) the mole fraction of gas D.

|  |
| --- |
| 1 – 0.870 = 0.130 |

(c) the partial pressure of gas D.

|  |
| --- |
| *p*(D) = *x*(D) × *P* = 0.130 × 644 = 83.7 Pa  OR  *p*(D) = *P* – *p*(C) = 644 – 560 = 84 Pa |

6. A 15.0 dm3 container contains an equilibrium mixture of 0.862 mol gas X and 1.83 mol gas Y. The temperature is 255 °C. The gas constant, *R*, is 8.314 J mol–1 K–1.

Calculate the partial pressures of X and Y in kPa.

|  |
| --- |
| The ideal gas equation needs to be used to determine the total pressure. First convert the units:  15.0 dm3 = 0.015 m3  255 °C = 528 K  *P* =  = = 7.878 × 105 Pa = 787.8 kPa  *x*(X) = 0.862 / 2.692 = 0.3202; *p*(X) = 0.3202 × 787.8 = 252 kPa  *x*(Y) = 1.83 / 2.692 = 0.6798; *p*(Y) = 0.6798 × 787.8 = 536 kPa |

### 

### Activity 2: *K*p expressions and units

**Answers**

1. For each of the following homogeneous gaseous equilibria, construct an expression for *K*p and state its units assuming pressures in Pa.

(a) N2O4(g) ⇌ 2NO2(g)

|  |
| --- |
| *K*p =  units: Pa |

(b) N2(g) + 3H2(g) ⇌ 2NH3(g)

|  |
| --- |
| *K*p =  units: Pa–2 |

(c) H2(g) + CO2(g) ⇌ H2O(g) + CO(g)

|  |
| --- |
| *K*p =  units: none |

2. For each of the following heterogeneous equilibria, construct an expression for *K*p and state its units assuming pressures in Pa.

(a) H2O(g) + C(s) ⇌ H2(g) + CO(g)

|  |
| --- |
| *K*p =  units: Pa |

(b) CaCO3(s) ⇌ CaO(s) + CO2(g)

|  |
| --- |
| *K*p = *p*(CO2) units: Pa |

### Activity 3: *K*p calculations with amounts at equilibrium

This activity – and the next, for calculations with initial amounts – advocates using a table method to work through the calculations. In this method, known information is written below each component of the equilibrium, and this then used to work out subsequent bits of information in turn – working towards the partial pressures. At that point, the value of *K*p can be calculated.

It can be helpful to model this method for learners, and then to tackle some problems collaboratively before having learners work individually. For example, write one of the equilibria from the question sheet – or from questions sourced elsewhere – on the board, and have learners take turns to write pieces of information below each component. The class can provide suggestions if anyone gets stuck, e.g. suggest a formula to use to calculate the mole fraction.

Next, learners can work in smaller groups. Each person in the group has a piece of paper or mini-whiteboard for one of the components in the reaction, and is responsible for writing down the successive pieces of information for that component. Again, other members of the group can provide suggestions if anyone gets stuck.

The final question in the activity does not provide equilibrium amounts, but a different selection of data. Encourage learners to think about what they can calculate using the data provided – they could perhaps draw the usual table, fill in the data provided it the question and then consider which gaps they can fill. If they remain stuck, you can provide the following hints:

* you can calculate the mole fraction of X (from its partial pressure and the total pressure)
* you can calculate the mole fraction of Z (from the mole fractions of X and Y)
* from here you can use the table method.

**Answers**

Calculate *K*p, including units, for the following equilibria using the data provided.

1. N2O4(g) ⇌ 2NO2(g)

*n*(N2O4) = 4.02 × 10–3 mol; *n*(NO2) = 1.77 × 10–2 mol; total pressure = 16.4 kPa

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| --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- |
| |  |  |  |  | | --- | --- | --- | --- | |  | N2O4 | ⇌ | NO2 | | amount at equilibrium, *n* / mol | 4.02 × 10–3 |  | 1.77 × 10–2 | | total amount: *N* = 4.02 × 10–3 + 1.77 × 10–2 = 2.172 × 10–2 mol | | | | | mole fraction, *x*  (= *n*/*N*) | 0.402 / 2.172  = 0.1851 |  | 1.77 / 2.172  = 0.8149 | | (check total of mole fractions is 1: 0.1851 + 0.8149 = 1) | | | | | partial pressure, *p*  (= *x* × *P*) / kPa | 0.1851 × 16.4  = 3.036 |  | 0.8149 × 16.4  = 13.36 |   *K*p = = = 58.8 kPa |

2. A(g) + 2B(g) ⇌ C(g) + D(g)

*n*(A) = 1.32 mol; *n*(B) = 0.661 mol; *n*(C) = 0.187 mol; *n*(D) = 0.535 mol;   
total pressure = 488 kPa

|  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |
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| |  |  |  |  |  | | --- | --- | --- | --- | --- | |  | A | B | C | D | | amount at equilibrium, *n* / mol | 1.32 | 0.661 | 0.187 | 0.535 | | total amount: *N* = 1.32 + 0.661 + 0.187 + 0.535 = 2.703 mol | | | | | | mole fraction, *x*  (= *n*/*N*) | 1.32 / 2.703  = 0.4883 | 0.661 / 2.703  = 0.2445 | 0.187 / 2.703  = 0.06918 | 0.535 / 2.703  = 0.1979 | | (check total of mole fractions: 0.4883 + 0.2445 + 0.06918 + 0.1979 = 1.00 to 3 sf) | | | | | | partial pressure, *p*  (= *x* × *P*) / kPa | 0.4883 × 488  = 238.3 | 0.2445 × 488  = 119.3 | 0.06918 × 488  = 33.76 | 0.1979 × 488  = 96.58 |   *K*p = = = 9.61 × 10–4 kPa–1 |

3. Ag2CO3(s) ⇌ Ag2O(s) + CO2(g)

Total pressure = 17.8 kPa

|  |
| --- |
| The solids are ignored when writing *K*p for a heterogeneous equilibrium. Therefore  *K*p = *p*(CO2) kPa  Since CO2 is the only gas present, the total pressure is also the partial pressure for CO2.  So, *K*p = 17.8 kPa. |

4. X(g) + 2Y(g) ⇌ 2Z(g)

Total pressure = 125 kPa; partial pressure of X = 33.0 kPa; mole fraction of Y = 0.37

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| --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- |
| Determining mole fractions:  *x*(X) = *p*(X) / *P* = 33.0 / 125 = 0.264  *x*(Z) = 1 – *x*(X) – *x*(Y) = 1 – 0.37 – 0.264 = 0.366  Now the table method can be applied, starting with the mole fractions:   |  |  |  |  | | --- | --- | --- | --- | |  | X | Y | Z | | mole fraction | 0.264 | 0.37 | 0.366 | | partial pressure, *p*  (= *x* × *P*) / kPa | 33.0 (from question) | 0.37 × 125  = 46.25 | 0.366 × 125  = 45.75 |   *K*p = = = 0.030 kPa–1 |

### Activity 4: *K*p calculations with initial amounts

**Answers**

In Activity 4, problems using initial amounts are introduced and therefore the table is expanded. In the answers given below, the amount of reactant lost and the amount of product gained are on separate rows, to make a separation between numbers that must be subtracted and those that must be added. You may prefer at some point to switch to, or even to start with, the ‘ICE’ method (Initial, Change, Equilibrium), in which the amounts lost and gained are shown on the same row. This is a little more compact.

1. 2NO(g) + O2(g) ⇌ 2NO2

1.6 mol NO is mixed with 1.4 mol O2 at constant temperature. At equilibrium, the mixture contains 0.40 mol NO and the total pressure is 1.8 MPa. Calculate *K*p at this temperature, including units.

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| |  |  |  |  | | --- | --- | --- | --- | |  | NO | O2 | NO2 | | initial amount / mol | 1.6 | 1.4 | 0 | | known equilibrium amount / mol | 0.40 |  |  | | apply reaction ratio 2 NO : 1 O2 : 2 NO2 | | | | | amount lost / mol | 1.20 | 1.2 / 2 = 0.60 | — | | amount gained / mol | — | — | 1.20 | | equilibrium amount / mol | 0.40 | 0.80 | 1.20 | | total amount at equilibrium: *N* = 0.40 + 0.80 + 1.20 = 2.40 mol | | | | | mole fraction, *x* (= *n*/*N*) | 0.40 / 2.40  = 0.1667 | 0.80 / 2.40  = 0.3333 | 1.20 / 2.40 = 0.5000 | | partial pressure, *p* (= *x* × *P*) / MPa | 0.1667 × 1.80 = 0.3001 | 0.3333 × 1.80 = 0.6000 | 0.5000 × 1.80 = 0.9000 |   *K*p =  = = 15 MPa–1 |

2. N2(g) + 3H2(g) ⇌ 2NH3(g)

15.0 mol N2 is mixed with 35.4 mol H2 at constant temperature. At equilibrium, the mixture contains 15.6 mol NH3 and the total pressure is 37.3 MPa. Calculate *K*p at this temperature, including units.

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| |  |  |  |  | | --- | --- | --- | --- | |  | N2 | H2 | NH3 | | initial amount / mol | 15.0 | 35.4 | 0 | | known equilibrium amount / mol |  |  | 15.6 | | apply reaction ratio 1 N2 : 3 H2 : 2 NH3 | | | | | amount gained / mol | — | — | 15.6 | | amount lost / mol | 15.6 / 2 = 7.80 | 3/2 × 15.6  = 23.4 | — | | equilibrium amount / mol | 7.20 | 12.0 | 15.6 | | total amount at equilibrium: *N* = 7.20 + 12.0 + 15.6 = 34.8 | | | | | mole fraction, *x* (= *n*/*N*) | 7.20 / 34.8  = 0.2069 | 12.0 / 34.8  = 0.3448 | 15.6 / 34.8 = 0.4483 | | partial pressure, *p* (= *x* × *P*) / MPa | 0.2069 × 37.3 = 7.717 | 0.3448 × 37.3 = 12.86 | 0.4483 × 37.3 = 16.72 |   *K*p =  =  = 1.70 × 10–2 MPa–2 |

3. H2O(g) + C(s) ⇌ H2(g) + CO(g)

22.8 mol H2O is mixed with 4.34 mol hydrogen over powdered carbon at 525 °C. At equilibrium, the mixture contains 8.06 mol carbon monoxide and the total pressure is 27.9 kPa. Calculate *K*p at this temperature, including units.

|  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |
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| Only the gases contribute to the *K*p calculation. No initial or equilibrium amounts are given for the solid carbon, so it can be ignored completely in the calculation.   |  |  |  |  | | --- | --- | --- | --- | |  | H2O | H2 | CO | | initial amount / mol | 22.8 | 4.34 | 0 | | known equilibrium amount / mol |  |  | 8.06 | | apply reaction ratio 1 H2O : 1 H2 : 1 CO | | | | | amount lost / mol | 8.06 | — | — | | amounts gained / mol | — | 8.06 | 8.06 | | equilibrium amount / mol | 14.74 | 12.40 | 8.06 | | total amount at equilibrium: *N* = 14.74 + 12.40 + 8.06 = 35.20 mol | | | | | mole fraction, *x* (= *n*/*N*) | 14.74 / 35.20  = 0.4188 | 12.40 / 35.20  = 0.3523 | 8.06 / 35.20 = 0.2290 | | partial pressure, *p* (= *x* × *P*) / kPa | 0.4188 × 27.9 = 11.68 | 0.3523 × 27.9 = 9.829 | 0.2290 × 27.9 = 6.389 |   *K*p =  =  = 5.38 kPa |

### Activity 5: Changing amounts of gases and *K*p

At A Level, learners need to learn to explain how an equilibrium constant (*K*c or *K*p) controls the position of equilibrium on changing concentration, pressure and temperature. This worksheet provides a route towards developing such explanations.

Learners could work through the worksheet in groups, with support where required, or the sheet could be used as the basis for a classroom discussion.

The sheet uses calculations to illustrate the general principle that is finally arrived at, and introduces the idea of the reaction quotient, *Q*. This achieves the following:

* Additional calculation practice.
* Increased engagement with what is physically happening in the system, in particular the idea of partial pressure.
* Introducing *Q* encourages thinking of two ‘steps’ in the process: first the effect of the change in amount of one of the gases, then the resulting chemical process that constitutes the shift in equilibrium. This may help to avoid some of the misconceptions that can arise in this area if the physical and chemical aspects to the situation become muddled.
* Overall, this enables a deeper understanding of the process than would be achieved by just learning the general abstract principle.

Note that the descriptions arrived at hold when changes to concentration are made at constant volume and temperature.

You may feel more comfortable following a qualitative route through the example rather than the quantitative route described. In that case, the worksheet / script can be adapted.

Following this example for a change in amount of one of the components, learners might follow a similar process to investigate what happens in response to a change in volume. For example, calculate the new partial pressures if the volume is halved, and investigate the difference between the reaction quotient and the *K*p value. It is recommended to work through a couple of examples to show how the direction of the shift depends on the reaction equation and thus the expression for *K*p.

A converse approach demonstrates the effect of changing temperature. Learners will already be familiar with the role of the enthalpy change of reaction in terms of how the equilibrium responds to a temperature change. It can then be shown how this leads to a change in the ratio of numerator to denominator, and thus a change in the value of *K*p (or *K*c).

**Answers**

For the reaction

2SO2(g) + O2(g) ⇌ 2SO3(g)

1. Write the expression for *K*p.

|  |
| --- |
| *K*p = |

At a certain temperature, the system contains 1 mol SO2, 2 mol O2 and 5 mol SO3. The total pressure is 200 kPa.

2. Calculate *K*p for this temperature.

|  |
| --- |
| *x*(SO2) = 1/8 = 0.125; *p*(SO2) = 0.125 × 200 = 25 kPa  *x*(O2) = 2/8 = 0.25; *p*(O2) = 0.25 × 200 = 50 kPa  *x*(SO3) = 5/8 = 0.625; *p*(SO3) = 0.625 × 200 = 125 kPa  *K*p =  = 0.5 kPa–1 |

Another 2 mol SO2 is added to the system. Assume the temperature and volume are kept constant.

3. (a) What happens to the pressure in the system?

|  |
| --- |
| The pressure increases. |

(b) Suggest the new value of the partial pressure of SO2 in the system *before* any shift in the position of equilibrium.

|  |
| --- |
| The 2 mol added has a partial pressure of 50 kPa (the 2 mol O2 originally present also had a partial pressure of 50 kPa; temperature, amount and volume are all constant). Added to the original partial pressure of 25 kPa gives 75 kPa. |

(c) What has happened to the partial pressures of O2 and SO3?

|  |
| --- |
| They remain the same – same amount of gas in constant volume at constant temperature. |

4. (a) Using the partial pressures from question 3 and the expression for *K*p, calculate *Q* for the system after addition of 2 mol SO2.

|  |
| --- |
| *Q* =  = 0.06 kPa–1 |

(b) Is the value of *Q* larger or smaller than the value of *K*p?

|  |
| --- |
| Smaller. |

5. Look back at your answer to question 4(b). Using the information above, in which direction will the equilibrium shift following the addition of SO2?

|  |
| --- |
| To the right / towards the product. The partial pressure of SO3 increases and the partial pressures of SO2 and O2 decrease until the ratio *K*p is restored. |

6. If you add more of a component of a gaseous equilibrium to the system, the partial pressure of that component ***increases***. The partial pressures of the other gases present ***remain the same***.

If the added gas is a reactant, the value of the ***denominator*** in the *K*p expression increases. The equilibrium shifts to the ***right*** to restore the ratio in the *K*p expression.

If the added gas is a product, the value of the ***numerator*** in the *K*p expression increases. The equilibrium shifts to the ***left*** to restore the ratio in the *K*p expression.

7. Now write similar descriptions to describe what happens if you **remove** one of the gases present at equilibrium.

|  |
| --- |
| Descriptions should express that removing a reactant reduces the value of the denominator, causing a shift to the left. Removing a product reduces the value of the numerator, causing a shift to the right. |

### Activity 6: The effect of conditions on the value of *K*p

This sheet contains further practice questions asking both how the value of *K*p and the position of equilibrium are affected by changes in conditions. The final two questions also include calculations for additional practice.

Again, these questions can be used as individual worksheets so that learners can practice formulating explanations, or as the basis for group or class discussions to help learners develop their understanding.

**Answers**

1. This question is about the reaction

N2(g) + 3H2(g) ⇌ 2NH3(g) Δ*H* = –92 kJ mol–1

(a) Write the expression for *K*p.

|  |
| --- |
| *K*p = |

(b) State and explain the effect on the value of *K*p when

(i) the partial pressure of H2 is increased.

|  |
| --- |
| No change. Increasing the partial pressure of one of the gases will result in a shift in the position of equilibrium, but does not affect the value of *K*p. |

(ii) an iron catalyst is added.

|  |
| --- |
| No change. A catalyst speeds up the forward and reverse reactions by the same proportion, and does not affect the position of equilibrium. |

(iii) the temperature is decreased.

|  |
| --- |
| The reaction is exothermic, so a decrease in temperature favours the forward reaction and shifts the position of equilibrium to the right. The partial pressure of NH3 increases and those of N2 and H2 decrease, so the value of *K*p increases. |

(c) Explain in terms of *K*p how this equilibrium will respond if some of the NH3 is removed from the system. Assume the volume and temperature are kept constant.

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| Removing NH3 causes a temporary decrease in the partial pressure of NH3. Because the value of *K*p remains constant, this decrease must be followed by the forward reaction to produce more NH3, increasing *p*(NH3). This proceeds until the ratio    is again equal to *K*p. |

2. This question is about the reaction

Cu2S(s) + O2(g) ⇌ 2Cu(s) + SO2(g) Δ*H* = –250 kJ mol–1

(a) Write the expression for *K*p.

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| *K*p = |

(b) State and explain the effect on the value of *K*p when

(i) a catalyst is added.

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| No change. A catalyst speeds up the forward and reverse reactions by the same proportion, and does not affect the position of equilibrium. |

(ii) the partial pressure of O2(g) is increased.

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| No change. Increasing the partial pressure of one of the gases will result in a shift in the position of equilibrium, but does not affect the value of *K*p. |

(c) Explain in terms of *K*p how the equilibrium will respond to the following changes. Assume the temperature remains constant.

(i) An increase in the mole fraction of O2.

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| Increasing the mole fraction of O2 causes a temporary increase in its partial pressure. Because the value of *K*p remains constant, this increase must be followed by the forward reaction removing some of the O2 – decreasing *p*(O2) – to produce more SO2. This proceeds until the ratio    is again equal to *K*p. |

(ii) Addition of Cu2S to the system.

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| There will be no change in the position of the equilibrium. Cu2S is a solid, and so does not feature in the expression for *K*p. |

(iii) An increase in the total pressure of the system.

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| The partial pressures of SO2 and O2 both increase proportionally to the increase in pressure. Therefore the ratio    remains the same, and there is no shift in the position of the equilibrium. |

3. This question is about the reaction

2NaHCO3(s) ⇌ Na2CO3(s) + H2O(g) + CO2(g) Δ*H* = +128 kJ mol–1

(a) Write the expression for *K*p.

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| *K*p = *p*(H2O) *p*(CO2) |

(b) State and explain the effect on the value of *K*p when

(i) the mole fraction of H2O(g) is decreased.

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| No change. Decreasing the mole fraction (and thus partial pressure) of one of the gases will result in a shift in the position of equilibrium, but does not affect the value of *K*p. |

(ii) the temperature is decreased.

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| The reaction is endothermic, so a decrease in temperature favours the reverse reaction and shifts the position of equilibrium to the left. The partial pressures of H2O and CO2 decrease, so the value of *K*p decreases. |

(c) Explain in terms of *K*p how the amount of Na2CO3 will be affected if H2O is removed from the system. Assume the temperature and volume remain constant.

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| Removing H2O from the system temporarily reduces its partial pressure. As *K*p remains constant, the equilibrium shifts to the right to increase the partial pressure of H2O­ and CO2 until the product *p*(H2O) × *p*(CO2) is again equal to *K*p. Because the equilibrium shifts to the right, the amount of Na2CO3 in the system also increases. |

4. This question is about the reaction

2X(g) + 3Y(g) ⇌ 4Z(g)

(a) Write the expression for *K*p for this equilibrium.

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| *K*p = |

(b) 0.50 mol X and 0.50 mol Y are mixed together at a certain temperature. At equilibrium, the system contains 0.40 mol Z. Calculate the equilibrium amounts of X and Y.

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| 0.20 mol X has reacted to form 0.40 mol Z, so 0.30 mol X remains. 0.30 mol Y has reacted to form 0.40 mol Z, so 0.20 mol Y remains. |

(c) Calculate the mole fractions of X, Y and Z at equilibrium.

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| *N* = 0.30 + 0.20 + 0.40 = 0.90 mol  *x*(X) = 0.30 / 0.90 = 0.333  *x*(Y) = 0.20 / 0.90 = 0.222  *x*(Z) = 0.40 / 0.90 = 0.444 |

(d) The total pressure at equilibrium is 150 kPa. Calculate the value of *K*p, including units.

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| *p*(X) = 0.333 × 150 = 50.0 kPa  *p*(Y) = 0.222 × 150 = 33.3 kPa  *p*(Z) = 0.444 × 150 = 66.7 kPa  *K*p = = 0.21 kPa–1 |

(e) Explain how *K*p and the equilibrium system will respond if the volume of the system is reduced. Assume the temperature is kept constant.

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| The value of *K*p is constant at constant temperature. If the volume is reduced, the partial pressures of all components increase proportionally. The expression for *K*p shows that the denominator increases by a greater proportion than the numerator. *K*p is constant, so the reaction must shift to reduce the size of the denominator and increase the size of the numerator, i.e. towards the product. |

5. This question is about the reaction

A(g) + 2B(s) ⇌ C(g) + 2D(g)

(a) Write the expression for *K*p for this equilibrium.

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| *K*p = |

(b) 0.80 mol A and 0.50 mol B are mixed together at a certain temperature. At equilibrium, the system contains 0.25 mol C. Calculate the equilibrium amounts of A, B and D.

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| 0.25 mol A has reacted to form 0.25 mol C, so 0.55 mol A remains. 0.50 mol B has reacted to form 0.25 mol C, so 0 mol B remains. Formation of 0.25 mol C means formation of 0.50 mol D. |

(c) Calculate the mole fractions of all gaseous components at equilibrium.

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| *N* = 0.55 + 0.25 + 0.50 = 1.30 mol  *x*(A) = 0.55 / 1.30 = 0.423  *x*(C) = 0.25 / 1.30 = 0.192  *x*(D) = 0.50 / 1.30 = 0.385 |

(d) The total pressure at equilibrium is 5.0 kPa. Calculate the value of *K*p, including units.

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| *p*(A) = 0.423 × 5.0 = 2.12 kPa  *p*(C) = 0.192 × 5.0 = 0.960 kPa  *p*(D) = 0.385 × 5.0 = 1.93 kPa  *K*p = = 1.7 kPa2 |

(e) The volume of the container is reduced while the temperature is kept constant. A deposit of solid B forms in the container. Explain why.

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| The value of *K*p is constant at constant temperature.  If the volume is reduced, the partial pressures of all components increase proportionally. The expression for *K*p shows that the numerator increases by a greater proportion than the denominator. *K*p is constant, so the reaction must shift to reduce the size of the numerator and increase the size of the denominator, i.e. towards the reactants. So, some B(s) re-forms. |

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# Topic Exploration Pack

## Equilibrium and pressure – Kp

## Learner activity

### Learner Activity 1: Mole fractions and partial pressures

This sheet gives you an introduction to the concept of mole fractions and partial pressures. You will apply these concepts later in calculations of equilibrium constants for gaseous equilibria (*K*p).

**The mole fraction**

The *mole fraction* of gas A (*x*(A)) is the amount of A in mol (*n*(A)) divided by the total amount of gas in the system (*N*). As an equation:



The total amount of gas includes the amount of A. So, for a system that contains the gases A, B and C

*N* = *n*(A) + *n*(B) + *n*(C)

The sum of all the mole fractions for the gases in the mixture must equal 1. So for the same system

*x*(A) + *x*(B) + *x*(C) = 1

**Example**

A mixture of gases contains 0.24 mol H2, 0.14 mol N2 and 5.51 mol NH3.

*N* = *n*(H2) + *n*(N2) + *n*(NH3) = 0.24 + 0.14 + 5.51 = 5.89 mol gas

Mole fractions:

*x*(H2) = 0.24 / 5.89 = 0.0407

*x*(N2) = 0.14 / 5.89 = 0.0238

*x*(NH3) = 5.51 / 5.89 = 0.935

Check: 0.0407 + 0.0238 + 0.935 = 1.00 – as expected, the mole fractions add up to 1.

**Questions**

1. An equilibrium mixture contains 0.50 mol I2, 0.75 mol H2 and 4.5 mol HI. Calculate the mole fraction of each gas.

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2. An equilibrium mixture contains two gases, A and B. The mole fraction of gas A is 0.46. Calculate the mole fraction of gas B.

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3. Explain why the mole fraction has no units.

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**Partial pressures**

The partial pressure of gas A (*p*(A)) is the mole fraction of gas A (*x*(A)) multiplied by the total pressure (*P*). As an equation:

*p*(A) = *x*(A) × *P*

The sum of all the partial pressures for gases in a system is equal to the total pressure. So for a system that contains the gases A, B and C:

*p*(A) + *p*(B) + *p*(C) = *P*

4. Look back at the equilibrium mixture described in question 1. The total pressure of this mixture is 115 kPa. Calculate the partial pressure of each gas in the mixture.

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5. An equilibrium mixture contains two gases, C and D. The mole fraction of C is 0.870. The partial pressure of C is 560 Pa. Calculate

(a) the total pressure.

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(b) the mole fraction of gas D.

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(c) the partial pressure of gas D.

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6. A 15.0 dm3 container contains an equilibrium mixture of 0.862 mol gas X and 1.83 mol gas Y. The temperature is 255 °C. The gas constant, *R*, is 8.314 J mol–1 K–1.

Calculate the partial pressures of X and Y in kPa.

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### Learner Activity 2: *K*p expressions and units

This sheet will allow you to practice constructing expressions for *K*p and determining the units.

Remember:

in a *homogeneous* equilibrium all the species are in the *same* state or phase

in a *heterogeneous* equilibrium species are in *different* states or phases.

For the general homogeneous gaseous equilibrium

*a*A(g) + *b*B(g) ⇌ *c*(C)(g) + *d*(D)(g)

the equilibrium constant, *K*p, is defined in terms of partial pressures as



Take care: the partial pressures above and below the line are *multiplied*, not added.

The units will be

(pressure unit, e.g. Pa)(*c*+*d*)–(*a*+*b*)

Note: if (*a* + *b*) = (*c* + *d*), *K*p has no units.

For heterogeneous equilibria, the *K*p expression includes the partial pressures of the gases as above. Any non-gaseous species in the reaction equation are ignored – they are not included in the expression for *K*p.

For example, for the equation

X(s) ⇌ Y(g) + Z(g)

*K*p = *p*(Y) × *p*(Z)

**Questions**

1. For each of the following homogeneous gaseous equilibria, construct an expression for *K*p and state its units assuming pressures in Pa.

(a) N2O4(g) ⇌ 2NO2(g)

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(b) N2(g) + 3H2(g) ⇌ 2NH3(g)

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(c) H2(g) + CO2(g) ⇌ H2O(g) + CO(g)

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2. For each of the following heterogeneous equilibria, construct an expression for *K*p and state its units assuming pressures in Pa.

(a) H2O(g) + C(s) ⇌ H2(g) + CO(g)

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(b) CaCO3(s) ⇌ CaO(s) + CO2(g)

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### Learner Activity 3: *K*p calculations with amounts at equilibrium

A general method for calculations where the equilibrium amounts are given is to set your calculation out underneath the reaction equation as a table. Write below each species:

* the equilibrium amount (in mol)
* the mole fraction
* the partial pressure.

Then construct the expression for *K*p, and substitute the values for the partial pressures. Complete the calculation, giving the final answer to the appropriate number of significant figures. Finally, state the units for *K*p.

**Questions**

Calculate *K*p, including units, for the following equilibria using the data provided.

1. N2O4(g) ⇌ 2NO2(g)

*n*(N2O4) = 4.02 × 10–3 mol; *n*(NO2) = 1.77 × 10–2 mol; total pressure = 16.4 kPa

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2. A(g) + 2B(g) ⇌ C(g) + D(g)

*n*(A) = 1.32 mol; *n*(B) = 0.661 mol; *n*(C) = 0.187 mol; *n*(D) = 0.535 mol; total pressure = 488 kPa

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3. Ag2CO3(s) ⇌ Ag2O(s) + CO2(g)

Total pressure = 17.8 kPa

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4. X(g) + 2Y(g) ⇌ 2Z(g)

Total pressure = 125 kPa; partial pressure of X = 33.0 kPa; mole fraction of Y = 0.37

(You cannot go straight into the table method with this question because only one of the equilibrium amounts is given. See if you can find a different way to tackle the question. You can ask your teacher for a hint if you get stuck.)

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### Learner Activity 4: *K*p calculations with initial amounts

Calculations where initial amounts are given can be confusing, and involve several steps. Again, constructing your answer using a table can be a good way of tackling the problem – making sure you methodically work through all the required steps. With initial amounts calculations, there are a few more things to think about. Below each species, write

* the initial amount (where known)
* the equilibrium amount (where known)
* amount lost / gained (this requires working out)
* the equilibrium amount, as deduced from the above
* mole fraction
* partial pressure.

Then, construct the *K*p equation and calculate the value as normal. Don’t forget the units!

**Questions**

1. 2NO(g) + O2(g) ⇌ 2NO2

1.6 mol NO is mixed with 1.4 mol O2 at constant temperature. At equilibrium, the mixture contains 0.40 mol NO and the total pressure is 1.8 MPa. Calculate *K*p at this temperature, including units.

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2. N2(g) + 3H2(g) ⇌ 2NH3(g)

15.0 mol N2 is mixed with 35.4 mol H2 at constant temperature. At equilibrium, the mixture contains 15.6 mol NH3 and the total pressure is 37.3 MPa. Calculate *K*p at this temperature, including units.

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3. H2O(g) + C(s) ⇌ H2(g) + CO(g)

22.8 mol H2O is mixed with 4.34 mol hydrogen over powdered carbon at 525 °C. At equilibrium, the mixture contains 8.06 mol carbon monoxide and the total pressure is 27.9 kPa. Calculate *K*p at this temperature, including units.

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### Learner Activity 5: Changing amounts of gases and *K*p

This sheet helps you to think about what happens when you add more of one of the components of an equilibrium to a system. You will work through some calculations to build up an explanation for why the equilibrium shifts in a particular direction. The explanation uses the expression and value for *K*p, rather than the more vague idea of ‘opposing change’.

The ideas in this sheet also apply to *K*c.

**Questions**

For the reaction

2SO2(g) + O2(g) ⇌ 2SO3(g)

1. Write the expression for *K*p.

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At a certain temperature, the system contains 1 mol SO2, 2 mol O2 and 5 mol SO3. The total pressure is 200 kPa.

2. Calculate *K*p for this temperature.

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Another 2 mol SO2 is added to the system. Assume the temperature and volume are kept constant.

3. (a) What happens to the pressure in the system?

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(b) Suggest the new value of the partial pressure of SO2 in the system *before* any shift in the position of equilibrium.

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(c) What has happened to the partial pressures of O2 and SO3?

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We will now introduce you to the reaction quotient, *Q*. *Q* can help you predict how an equilibrium will shift if you disturb it, for example by adding more of one of the components. *Q* is calculated using the same expression as *K*p, but you insert the partial pressures immediately after the equilibrium has been disturbed – before the equilibrium has settled to accommodate the change.

4. (a) Using the partial pressures from question 3 and the expression for *K*p, calculate *Q* for the system after addition of 2 mol SO2.

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(b) Is the value of *Q* larger or smaller than the value of *K*p?

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We know that *K*p only changes if there is a change in temperature. So, if following a change to the equilibrium the value of *Q* is not the same as the value of *K*p­, the equilibrium must shift. The equilibrium shifts until the partial pressures of all the components again produce the value of *K*p.

If *Q* is larger than *K*p, the numerator of the *K*p expression is too large relative to the denominator. So the partial pressure of the product needs to reduce, and the partial pressure of the reactants needs to increase, until the ratio is again equal to *K*p. That means the equilibrium shifts to the left, towards the reactants.

The opposite argument is true if *Q* is smaller than *K*p.

5. Look back at your answer to question 4(b). Using the information above, in which direction will the equilibrium shift following the addition of SO2?

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You need to be able to describe shifts in the position of an equilibrium in terms of *K*p. Complete the following description of what happens if you add more of one of the components of a gaseous equilibrium.

6. If you add more of a component of a gaseous equilibrium to the system, the partial pressure of that component \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_. The partial pressures of the other gases present \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_.

If the added gas is a reactant, the value of the \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ in the *K*p expression increases. The equilibrium shifts to the \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ to restore the ratio in the *K*p expression.

If the added gas is a product, the value of the \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ in the *K*p expression increases. The equilibrium shifts to the \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ to restore the ratio in the *K*p expression.

7. Now write similar descriptions to describe what happens if you **remove** one of the gases present at equilibrium.

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### Learner Activity 6: The effect of conditions on the value of *K*p

This sheet contains questions to help you review how changing certain factors does or doesn’t affect the value of *K*p. There are also some calculations at the end to give you some more practice.

**Questions**

1. This question is about the reaction

N2(g) + 3H2(g) ⇌ 2NH3(g) Δ*H* = –92 kJ mol–1

(a) Write the expression for *K*p.

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(b) State and explain the effect on the value of *K*p when

(i) the partial pressure of H2 is increased.

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(ii) an iron catalyst is added.

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(iii) the temperature is decreased.

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(c) Explain in terms of *K*p how this equilibrium will respond if some of the NH3 is removed from the system. Assume the volume and temperature are kept constant.

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2. This question is about the reaction

Cu2S(s) + O2(g) ⇌ 2Cu(s) + SO2(g) Δ*H* = –250 kJ mol–1

(a) Write the expression for *K*p.

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(b) State and explain the effect on the value of *K*p when

(i) a catalyst is added.

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(ii) the partial pressure of O2(g) is increased.

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(c) Explain in terms of *K*p how the equilibrium will respond to the following changes. Assume the temperature remains constant.

(i) An increase in the mole fraction of O2.

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(ii) Addition of Cu2S to the system.

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(iii) An increase in the total pressure of the system.

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3. This question is about the reaction

2NaHCO3(s) ⇌ Na2CO3(s) + H2O(g) + CO2(g) Δ*H* = +128 kJ mol–1

(a) Write the expression for *K*p.

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(b) State and explain the effect on the value of *K*p when

(i) the mole fraction of H2O(g) is decreased.

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(ii) the temperature is decreased.

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(c) Explain in terms of *K*p how the amount of Na2CO3 will be affected if H2O is removed from the system. Assume the temperature and volume remain constant.

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4. This question is about the reaction

2X(g) + 3Y(g) ⇌ 4Z(g)

(a) Write the expression for *K*p for this equilibrium.

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(b) 0.50 mol X and 0.50 mol Y are mixed together at a certain temperature. At equilibrium, the system contains 0.40 mol Z. Calculate the equilibrium amounts of X and Y.

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(c) Calculate the mole fractions of X, Y and Z at equilibrium.

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(d) The total pressure at equilibrium is 150 kPa. Calculate the value of *K*p, including units.

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(e) Explain how *K*p and the equilibrium system will respond if the volume of the system is reduced. Assume the temperature is kept constant.

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5. This question is about the reaction

A(g) + 2B(s) ⇌ C(g) + 2D(g)

(a) Write the expression for *K*p for this equilibrium.

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(b) 0.80 mol A and 0.50 mol B are mixed together at a certain temperature. At equilibrium, the system contains 0.25 mol C. Calculate the equilibrium amounts of A, B and D.

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(c) Calculate the mole fractions of all gaseous components at equilibrium.

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(d) The total pressure at equilibrium is 5.0 kPa. Calculate the value of *K*p, including units.

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(e) The volume of the container is reduced while the temperature is kept constant. A deposit of solid B forms in the container. Explain why.

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1. Quílez, J. (2004), *Chem. Ed. Res. Pract.*, 5, 281–300 [↑](#footnote-ref-1)
2. Cheung, D. (2009), *J. Chem. Ed.*, 86, 514–518 [↑](#footnote-ref-2)