

CHEMICAL EQUILIBRIA: GASES

413 Team 1

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MOLE FRACTIONS AND PARTIAL PRESSURE

In a mixture of chemicals (A, B, C, etc.), the **mole fraction** of chemical A, $x(A)$, is defined as:

$$x(A) = \frac{\text{no. moles of A}}{\text{total no. moles}}$$

If the chemicals are gases, 'no. moles' could be replaced by 'volume':

$$x(A) = \frac{\text{volume of A}}{\text{total volume}}$$

The **partial pressure** of a gas in a mixture is the proportion of the total pressure due to that gas only. Partial pressure of A, $p(A)$, is related to the mole fraction of A and the total pressure, P :

$$p(A) = x(A) \times P$$

If any of the reactants or products are **not** gases, they must be left out of the K_p expression (solids and liquids do not have a partial pressure)

HOW TEMPERATURE AFFECTS HOMOGENEOUS GAS-PHASE EQUILIBRIA

Look back at **Chemical Equilibrium Part 1** (sheet B) and review *Le Chatelier's principle*. It predicts that if you increase the temperature of a system at equilibrium, the position of equilibrium will move in whichever direction is an **endothermic** reaction, and if you decrease the temperature, the position of equilibrium moves whichever way is **exothermic**.

The reaction of hydrogen with iodine is exothermic: $\text{H}_{2(g)} + \text{I}_{2(g)} \rightleftharpoons 2\text{HI}_{(g)}$ $\Delta H = -ve$ and $K_p = \frac{p(\text{HI})^2}{p(\text{H}_2)p(\text{I}_2)}$

If this reaction is at equilibrium and then is heated up, the position of equilibrium will move **left**. How will this affect the value of K_p ?

As there will be more moles of H_2 and I_2 and fewer moles of HI , in the new equilibrium mixture, the new value of K_p will be lower.

In general, for **exothermic** reactions:

As temperature falls, K_p increases



As temperature rises, K_p decreases

For **endothermic** reactions:

As temperature falls, K_p decreases

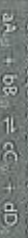


As temperature rises, K_p increases

EQUILIBRIUM CONSTANT, K_p

In equilibria involving gases, the equilibrium constant can be written in terms of the partial pressures of the gaseous reactants and products (instead of using concentrations).

For an equilibrium with a **balanced** equation like this:



the equilibrium constant, K_p can be written as:

$$K_p = \frac{p(\text{C})^c p(\text{D})^d}{p(\text{A})^a p(\text{B})^b}$$

HOW PRESSURE AFFECTS HOMOGENEOUS GAS-PHASE EQUILIBRIA

Carbon monoxide reacts with oxygen in an equilibrium reaction as shown here: $2\text{CO}_{(g)} + \text{O}_{2(g)} \rightleftharpoons 2\text{CO}_{2(g)}$ In the forward reaction the number of moles of gas decreases (the balanced equation shows 3 moles of reactants and 2 moles of product).

According to *Le Chatelier's principle*, when the pressure increases, the position of equilibrium will move to the right (where the fewest moles of gas are). This will cause $x(\text{CO})$ and $x(\text{O}_2)$ to decrease while $x(\text{CO}_2)$ will increase.

This does not mean that K_p will change, though; $p(\text{CO})$, $p(\text{O}_2)$ and $p(\text{CO}_2)$ change in such a way that the fraction $\frac{p(\text{CO}_2)^2}{p(\text{CO})^2 p(\text{O}_2)}$ has the same value as before pressure was increased.

REMEMBER THIS

Increasing temperature moves the position of equilibrium to the **right** and increases K_p for **endothermic** reactions. Increasing temperature moves the position of equilibrium to the **left** and decreases K_p for **exothermic** reactions. Changes in pressure or the addition of catalysts do not change the numerical value of K_p .