Reaction Quotient

Q

Reaction Quotient

How would we know if equilibrium has been reached? Or if it will be reach?

The **reaction quotient** , **Q**, or **trial KC**, enables us to determine this information.

The reaction quotient is determined by using the equilibrium law or mass action expression and substituting either initial concentrations or those determined during experimental trials.

Reaction Quotient

To determine which reaction is favoured and in which direction the system is moving, Q is compared to KC:

If Q = K, the system is at equilibrium. The forward and reverse rates are equal and the reactant and product concentrations remain constant.

If Q > K, the system is NOT at equilibrium. There is too much product, so the reverse reaction is favoured to bring the reactant-product ratio to equal K

If Q < K, the system is NOT at equilibrium. The concentration of reactants is too large, so the forward reaction is favoured.

Example

Reaction quotient is based on the initial concentrations, while equilibrium constant is based on the concentration AT equilibrium!

$$N_2(g) + O_2(g) \rightleftharpoons 2\ NO(g)$$

It was found that 8.50 moles of nitrogen, 11.0 moles of oxygen and 2.20 moles of nitrogen monoxide were in a 5.00 L container. If the equilibrium constant is 0.035, what are the following concentrations at equilibrium?

If not, which reaction is favoured and which concentrations are increasing and which are decreasing?

Reaction Quotient

Reaction quotient is based on the initial concentrations, while equilibrium constant is based on the concentration AT equilibrium!

$$N_2(g) + O_2(g) \rightleftharpoons 2\ NO(g)$$

Step 1: Find concentrations for all

$$C_{N_2} = \frac{8.50\ \text{moles}}{5.00\ \text{L}} = 1.70\ \text{mol/L}$$

$$C_{O_2} = \frac{11.0\ \text{moles}}{5.00\ \text{L}} = 2.20\ \text{mol/L}$$

$$C_{NO} = \frac{2.20\ \text{moles}}{5.00\ \text{L}} = 0.440\ \text{mol/L}$$

Reaction Quotient

Reaction quotient is based on the initial concentrations, while equilibrium constant is based on the concentration AT equilibrium!

$$N_2(g) + O_2(g) \rightleftharpoons 2\,NO(g)$$

Step 2: Substitute the concentrations in the equilibrium law and replace K with Q.

$$Q = \frac{[NO]^2}{[N_2][O_2]} = \frac{(0.440)^2}{(1.70)(2.20)} = 0.0518$$

Reaction Quotient

Reaction quotient is based on the initial concentrations, while equilibrium constant is based on the concentration AT equilibrium!

$$N_2(g) + O_2(g) \rightleftharpoons 2\,NO(g)$$

Step 3: Compare the Q with K, are they equal? Smaller? Larger?

In this example, the trial K or reaction quotient, Q, is greater than the value of K.

This means:

❖the system is not at equilibrium since K does not equal Q since $Q > K$

❖the reverse reaction is favoured and

❖the NO concentration is decreasing and the oxygen and nitrogen concentrations are increasing to reach equilibrium.

2nd Example

Example 2: For the following imaginary equilibrium system the value of K is 0.222

$$2A(g) + B(g) \leftrightarrow 3C(g)$$

1.50 moles of A and 3.20 moles of B are placed into a 1.0 L container and allowed to react. After several minutes, a sample is taken and found to contain 1.00 moles of A. Is the system at equilibrium? Which reaction rate is fastest? Which concentrations are increasing?

To answer part 1: We must determine the reaction quotient and compare it to K.

We can use the ICE Table

	2 A(g)	+ B (g)	$\rightleftharpoons$ 3 C(g)
I	1.50	3.20	0
C			
E	1.00	?	?

2nd Example

Example 2: For the following imaginary equilibrium system the value of K is 0.22

$$2A(g) + B\,(g) \leftrightarrow 3\,C(g)$$

1.50 moles of A and 3.20 moles of B are placed into a 1.0 L container and allowed to react. After several minutes, a sample is taken and found to contain 1.00 moles of A. Is the system at equilibrium? Which reaction rate is fastest? Which concentrations are increasing?

We can solve for the concentration when it reacts
1.50 (Initial) – 1.00 (After several Minutes) = 0.50

Get from Stoichiometry
0.50 of A * (1B/2A) = 0.25

We can use the ICE Table

	2 A(g)	+ B (g)	$\rightleftharpoons$ 3 C(g)
I	1.50	3.20	0
C	**0.50**	**0.25**	**0.75**
E	1.00	?	?

2nd Example

Example 2: For the following imaginary equilibrium system the value of K is 0.22

$2A(g) + B(g) \leftrightarrow 3\,C(g)$

1.50 moles of A and 3.20 moles of B are placed into a 1.0 L container and allowed to react. After several minutes, a sample is taken and found to contain 1.00 moles of A. Is the system at equilibrium? Which reaction rate is fastest? Which concentrations are increasing?

Subtract them to give you E

	2 A(g)	+ B (g)	$\rightleftharpoons$ 3 C(g)
I	1.50	3.20	0
C	–0.50	$-0.50\left(\frac{1B}{2A}\right) = 0.25$	$+0.50\left(\frac{3C}{2A}\right) = 0.75$
E	1.00	2.95	0.75

2nd Example

Example 2: For the following imaginary equilibrium system the value of K is 0.22

$$2A(g) + B(g) \leftrightarrow 3C(g)$$

1.50 moles of A and 3.20 moles of B are placed into a 1.0 L container and allowed to react. After several minutes, a sample is taken and found to contain 1.00 moles of A. Is the system at equilibrium? Which reaction rate is fastest? Which concentrations are increasing?

Plug the new values in. The values we are plugging in are NOT the "K" or equilibrium after several minutes. We are finding the QUOTIENT, which means it is the concentration at any time.

2nd Example

Example 2: For the following imaginary equilibrium system the value of K is 0.22

$$2A(g) + B(g) \leftrightarrow 3C(g)$$

1.50 moles of A and 3.20 moles of B are placed into a 1.0 L container and allowed to react. After several minutes, a sample is taken and found to contain 1.00 moles of A. Is the system at equilibrium? Which reaction rate is fastest? Which concentrations are increasing?

From the concentration

$$Q = \frac{[C]^3}{[A]^2[B]}$$

$$= \frac{(0.75)^3}{(1.00)^2(2.95)}$$

$$= 0.14$$

2nd Example

Example 2: For the following imaginary equilibrium system the value of K is 0.22

$$2A(g) + B(g) \leftarrow\rightarrow 3C(g)$$

1.50 moles of A and 3.20 moles of B are placed into a 1.0 L container and allowed to react. After several minutes, a sample is taken and found to contain 1.00 moles of A. Is the system at equilibrium? Which reaction rate is fastest? Which concentrations are increasing?

Compare the Q with K.

0.14 < 0.22, Q < K . If Q is < K then that means the system has not reached equilibrium. The forward reaction is faster than the reverse reaction. The concentration of products is increasing.

- For the reaction below, K = 25.0:

A(s) + 2 B(g) ←→ 2 C(g)

If [A] = 12.0 mol/L [B] = 2.0 mol/L [C] = 30.0 mol/L

a) Is the system is at equilibrium.

b) Which reaction is faster (favoured), forward or reverse?

c) Which concentrations are increasing or decreasing?

2nd Example

There exists an equilibrium if
5.0 moles of CO_2, 5.0 moles of CO and 0.20 moles of O_2 are in a 2.0 L container at 562°C.

Find KC for the reaction

$$2\ CO(g) + O_2(g)\ \ 2\ CO_2(g)$$

Would the system be at equilibrium if
$[CO_2]$ = 15.8 mol/L, [CO] = 10.0 mol/L and $[O_2]$ = 0.25 mol/L?

If not, which reaction is favoured?

- For the reaction below, K = 16.0

  $2\ SO_2(g) + O_2(g) \leftrightarrow 2\ SO_3(g)$
- Initially, [SO_2] = 5.0 mol/L, [O_2] = 10.0 mol/L and [SO_3] = 0.
- After two hours [O_2] = 7.9 mol/L. Is the system at equilibrium?
- If not, which substances are increasing and which are decreasing?