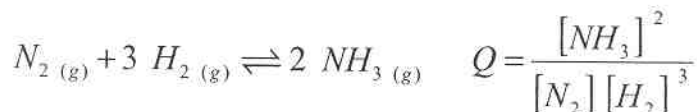


To determine the shift, we use the **reaction quotient,  $Q$** . The reaction quotient is obtained by applying the law of mass action using initial concentrations instead of equilibrium concentrations.



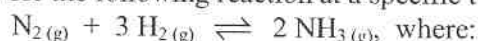
To determine in which direction a reaction will shift to reach equilibrium, we compare the values of  $Q$  and  $K$ .

- $Q$  is equal to  $K$ . The system is in equilibrium.
- $Q$  is greater than  $K$ . The system will shift to the left (toward the reactants)
- $Q$  is less than  $K$ . The system will shift to the right (toward the products)

$Q = K$ ; at equilibrium  
 $Q > K$ ; shift to the left  
 $Q < K$ ; shift to the right

Example:

Calculate the reaction quotient ( $Q$ ) for the following reaction at a specific temperature:



$$[N_2] = 4.6 \times 10^{-1} \text{ mol L}^{-1}$$

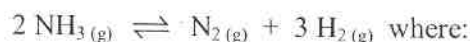
$$[H_2] = 2.2 \times 10^{-1} \text{ mol L}^{-1}$$

$$[NH_3] = 6.3 \times 10^{-1} \text{ mol L}^{-1}$$

$$Q = \frac{[NH_3]^2}{[N_2][H_2]^3} = \frac{(6.3 \times 10^{-1})^2}{(4.6 \times 10^{-1})(2.2 \times 10^{-1})^3} = 81.0$$

Example #2:

Calculate the reaction quotient ( $Q$ ) for the following reaction at a specific temperature:



$$[N_2] = 4.6 \times 10^{-1} \text{ mol L}^{-1}$$

$$[H_2] = 2.2 \times 10^{-1} \text{ mol L}^{-1}$$

$$[NH_3] = 6.3 \times 10^{-1} \text{ mol L}^{-1}$$

$$Q = \frac{[N_2][H_2]^3}{[NH_3]^2} = \frac{(4.6 \times 10^{-1})(2.2 \times 10^{-1})^3}{(6.3 \times 10^{-1})^2} = 0.012$$

We can draw a conclusion from the two examples (hopefully). If a reaction proceeds in the reverse direction under the same conditions the new equilibrium expression can be written as:



$$Q' = \frac{[A]^a [B]^b}{[C]^c [D]^d} = \frac{1}{Q}$$

### Equilibrium Position:

Again, the equilibrium concentrations will not always be the same, but the equilibrium constant ( $K$ ) remains the same at a given temperature. Using the data from the first example (front of the page) the conditions at equilibrium



Initial concentrations	Equilibrium concentrations
$[N_2]_0 = 4.7 \times 10^{-1} \text{ mol L}^{-1}$	$[N_2] = 4.6 \times 10^{-1} \text{ mol L}^{-1}$
$[H_2]_0 = 2.9 \times 10^{-1} \text{ mol L}^{-1}$	$[H_2] = 2.2 \times 10^{-1} \text{ mol L}^{-1}$
$[NH_3]_0 = 0 \text{ mol L}^{-1}$	$[NH_3] = 6.3 \times 10^{-1} \text{ mol L}^{-1}$

Each set of equilibrium concentrations is called an equilibrium position (there can be an infinite number of positions) but only one equilibrium constant ( $K$ ) for a given temperature. Look at the graph on the front side and pick any point in time (before equilibrium) that represents an equilibrium position.