# Equilibrium

**Unit 13B:** ICE Problems

### Skills:

* solve changing equilibrium problems using an ICE chart

ICE chart: a system for solving problems that involve a change in equilibrium conditions. The letters ICE stand for:

**Initial**: the amount of each component before the change

**Change**: the change to each component. Usually ± some multiple of *x*

**Equilibrium**: the amount of each component after the change. An equation created from the I and C terms.

## Strategy

1. Determine initial concentrations (in mol/L) or partial pressures of each component.
2. Create an ICE chart and list the initial concentrations.
3. Determine which direction the reaction will proceed. (This may require you to calculate the reaction quotient Q.)
4. Assign the variable “*x*” to the change in one species.
5. Calculate the changes to the other species in terms of *x*.
6. Express the equilibrium concentration of each species as an expression in *x*.
7. Substitute these expressions into the equilibrium expression and solve for *x*.

### Sample Problem

Given the reaction:

CO2 (g) + H2 (g)  CO (g) + H2O (g) *Kp* = 0.64 at 900 K

If you start with CO2 and H2 gas, each with a partial pressure of 1.00 atm, what are the partial pressures of the gases in the system at equilibrium.

1. Initial pressures:









1. ICE table:

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
|  | CO2 | H2 | CO | H2O |
| initial | 1.00 | 1.00 | 0 | 0 |
| change |  |  |  |  |
| equilibrium |  |  |  |  |

1. Because there is no CO or H2O present initially, the reaction will proceed to the right.
2. Because the coefficients are all 1, *x* = the decrease in  and  and also the increase in  and .
3. & 6. Our ICE table now looks like this:

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
|  | CO2 | H2 | CO | H2O |
| initial | 1.00 | 1.00 | 0 | 0 |
| change | −*x* | −*x* | +*x* | +*x* |
| equilibrium | 1.00 − *x* | 1.00 − *x* | *x* | *x* |

1. 

### Simplifying Assumption for Small Values of K

If your K value is very small (less than 10−5), the value of *x* will be extremely small, and anything that decreases by *x* will be approximately equal to the original value (*e.g.,* if *x* is small enough, then 0.25 − *x* ≈ 0.25).

For example:

A 200.0 ℓ container initially contains 45.3 mol N2 and 153.5 mol H2 at 725K. The following reaction occurs:

N2 (g) + 3 H2 (g)   2 NH3 (g) *Kc* = 5.3 × 10−5 at 725 K

Assuming ideal gas behavior, calculate the number of moles of NH3 present at equilibrium.

1. Initial concentrations are:



1. ICE table:

|  |  |  |  |
| --- | --- | --- | --- |
|  | N2 | H2 | NH3 |
| Initial | 0.227 | 0.767 | 0 |
| change |  |  |  |
| equilibrium |  |  |  |

1. Because there is no NH3 present initially, the reaction will proceed to the right.
2. Because the coefficient of N2 is 1, let *x* = the decrease in [N2]. This way, the other coefficients will be larger than 1 (instead of fractions).
3. & 6. Our ICE table now looks like this:

|  |  |  |  |
| --- | --- | --- | --- |
|  | N2 | H2 | NH3 |
| initial | 0.227 | 0.767 | 0 |
| change | −*x* | −3*x* | +2*x* |
| equilibrium | 0.227 − *x* | 0.767 − 3*x* | 2*x* |
| approximate equilibrium | 0.227 | 0.767 | 2*x* |

(Because *Kc* is very small, *x* will be very small, and we can approximate 0.227 − *x* ≈ 0.227 and 0.227 − 3*x* ≈ 0.767.)

1. 
2. [NH3] = 2*x* = 0.00233 M

Now answer the original question: 

Notice that both of those example problems included a specific scenario where you were initiating the reaction and therefore you had no products present at all in the system. Because of this, we could make the assumption that the reaction would proceed to the right (towards products) in at least a minuscule amount.

You must also be able to apply ice problems to scenarios where the reaction has been progressing for some time. In order to know which direction the reaction will proceed in, and therefore whether or not to add or subtract “x” in your ice table, you must determine which side of equilibrium you are on by calculating your Q value.

If Q and Keq = [products]/[reactants], then…

If Q = K: The reaction is at equilibrium, no change in concentration will occur

If Q > K: There are more products present than there will be at equilibrium.

* + The reaction will shift to the left (products into reactants)
  + Subtract “x” from products, add it to reactants

If Q < K: There are more reactants present than there will be at equilibrium.

* + The reaction will shift to the right (reactants into products)
  + Subtract “x” from reactants, add it to products

For example:

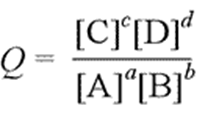
The same 200.0 ℓ container now contains 23.2 mol N2, 100.3 mol H2, and30.1 mol NH3 at 1025K. The following reaction occurs:

N2 (g) + 3 H2 (g)   2 NH3 (g) *Kc* = 8.4 × 10−3 at 1025 K

Assuming ideal gas behavior:

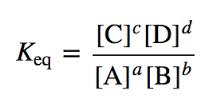
* Is the system at equilibrium?
* In which direction will the reaction proceed?
* How many moles will each reactant/product change by?

1. Figure out whether or not the system reached equilibrium (K = Q?)



Q =

1. Based on the Q value, which direction will the reaction move towards?
2. Determine the equilibrium concentration values and compare it to the current values to predict how much each will change.



Keq = 8.4 × 10−3 at 1025 K

8.4 × 10−3 =

X =

[N2] = [H2] = [NH3] =

Convert into moles and compare to what you started with: