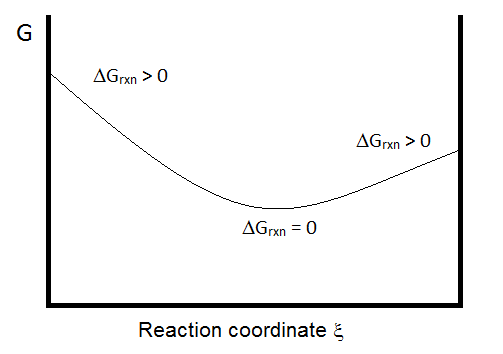
# Chapter 9: Chemical Equilibrium

The small is great, the great is small; all is in equilibrium in necessity...

- Victor Hugo in “Les Miserables”

As was discussed in Chapter 6, the natural tendency of chemical systems is to seek a state of minimum Gibbs function. Once the minimum is achieved, movement in any chemical direction will not be spontaneous. It is at this point that the system achieves a state of equilibrium.



From the diagram above, it should be clear that the direction of spontaneous change is determined by . If the slope of the curve is negative, the reaction will favor a shift toward products. And if it is positive, the reaction will favor a shift toward reactants. This is a non-trivial point, as it underscores the importance of the composition of the reaction mixture in the determination of the direction of the reaction.

## Chemical Potential

Equilibrium can be understood as accruing at the composition of a reaction mixture at which the aggregate chemical potential of the products is equal to that of the reactants. Consider the simple reaction

A(g) ↔ B(g)

The criterion for equilibrium will be

If the gases behave ideally, the chemical potentials can be described in terms of the mole fractions of A and B

where Dalton’s Law has been used to express the mole fractions.

The above expression can be simplified by putting all of the terms on the left and the terms on the right.

Combining the terms and recognizing that is –Go for the reaction, one obtains

And since pA/pB is Kp for this reaction (assuming perfectly ideal behavior), one can write

Another way to achieve this result is to consider the Gibbs function change for a reaction mixture in terms of the **reaction quotient**. The reaction quotient can be expressed as

where i are the stoichiometric coefficients for the products, and j are those for the reactants. Or if the stoichiometric coefficients are defined by expressing the reaction as a sum

Where Xi refers to one of the species in the reaction, and i is then the stoichiometric coefficient for that species, it is clear that i will be negative for a reactant (since its concentration or partial pressure will reduce as the reaction moves forward) and positive for a product (since the concentration or partial pressure will be increasing.) If the stoichiometric coefficients are expressed in this way, the expression for the reaction quotient becomes

Using this expression, the Gibbs function change for the system can be calculated from

And since at equilibrium

and

It is evident that

It is in this simple way that Kp and Go are related.

It is also of value to note that the criterion for a spontaneous chemical process is that Grxn < 0, rather than Grxno, as is stated in many texts! Recall that Grxno is a function of all of the reactants and products being in their standard states of unit fugacity or activity. However, the direction of spontaneous change for a chemical reaction is dependent on the composition of the reaction mixture. Similarly, the magnitude of the equilibrium constant is insufficient to determine whether a reaction will spontaneously form reactants or products, as the direction the reaction will shift is also a function of not just the equilibrium constant, but also the composition of the reaction mixture!

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| **Example**:  Based on the data below at 298 K, calculate the value of the equilibrium constant (Kp) for the reaction  2 NO(g) + O2(g) ↔ 2 NO2(g)   |  |  |  | | --- | --- | --- | |  | NO(g) | NO2(g) | | (kJ/mol) | 86.55 | 51.53 |   **Solution**:  First calculate the value of from the data.  And now use the value to calculate Kp using .  *Note*: as expected for a reaction with a very large negative , the equilibrium constant is very large, favoring the formation of the products. |

### Activities and Fugacities

To this point, we have mostly ignored deviations from ideal behavior. But it should be noted that thermodynamic equilibrium constants are not expressed in terms of concentrations or pressures, but rather in terms of activities and fugacities (both being discussed in Chapter 7). Based on these quantities,

and

And since activities and fugacities are unitless, thermodynamic equilibrium constants are unitless as well.

Further, it can be noted that the activities of solids and pure liquids are unity (assuming ideal behavior) since they are in their standard states at the given temperature. As such, these species never change the magnitude of the equilibrium constant and are generally omitted from the equilibrium constant expression.

### Kp and Kc

Oftentimes it is desirable to express the equilibrium constant in terms of concentrations (or activities for systems that deviate from ideal behavior.) In order to make this conversion, the relationship between pressure and concentration from the ideal gas law can be used.

And noting that the concentration is given by , the expression for the equilibrium constant becomes

And since for a given temperature, RT is a constant and can be factored out of the expression, leaving

This conversion works for reactions in which all reactants and products are in the gas phase. Care must be used when applying this relationship to heterogeneous equilibria.

## Pressure Dependence of Kp

Since the equilibrium constant Kp is a function of which is defined for a specific composition (all reactants in their standard states and at unit pressure (or fugacity), Changes in pressure have no effect on equilibrium constants for a fixed temperature. However, changes in pressure can have profound effects on the compositions of equilibrium mixtures.

To demonstrate the relationship, one must recall Dalton’s law of partial pressures. According to this relationship, the partial pressure of a component of a gas-phase mixture can be expressed

It is the combination of mole fractions that describes the composition of the equilibrium mixture.

Substituting the above expression into the expression for Kp yields

This expression can be factored into two pieces – one containing the mole fractions and thus describing the composition, and one containing the total pressure.

The second factor is a constant for a given total pressure. If the first term is given the symbol Kx, the expression becomes

In this expression, Kx has the same form as an equilibrium constant

but is not itself a constant. The value of Kx will vary with varying composition, and will *need* to vary with varying total pressure (in most cases) in order to maintain a constant value of Kp.

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| **Example**:  Consider the following reaction at equilibrium.  A(g) + 2 B(g) ↔ C(g) + D(g)  In which direction will the equilibrium shift if the volume of the reaction vessel is decreased?  **Solution**:  A decrease in the volume will lead to an increase in total pressure. Since the equilibrium constant can be expressed as  An increase in pressure will lead to an increase in Kx in order to maintain a constant value of Kp. So the reaction will shift to form more of the products C and D.  *Note*: This should make some sense, since a shift to the side of the reaction with fewer moles of gas will lower the total pressure of the reaction mixture, and thus relieving the stress introduced by increasing the pressure. This is exactly what is expected according to Le Chatlier’s principle. |

It should be noted that there are several ways one can affect the total pressure of a gas-phase equilibrium. These include the introduction or removal of reactants or products (perhaps through condensation or some other physical process), a change in volume of the reaction vessel, or the introduction of an inert gas that does not participate in the reaction itself. (Changes in the temperature will be discussed in a later section.) The principle of Le Chatlier can be used as a guide to predict how the equilibrium composition will respond to a change in pressure.

**Le Chatlier’s principle**: When a stress is introduced to a system at equilibrium, the system will adjust so as to reduce the stress.

Le Chatlier’s principle is fairly clear on how to think about the addition or removal of reactants or products. For example, the addition of a reactant will cause the system to shift to reduce the partial pressure of the reactant. It can do this by forming more products.

An important exception to the rule that increasing the total pressure will cause a shift in the reaction favoring the side with fewer moles of gas occurs when the total pressure is increased by introducing an inert gas to the mixture. The reason is that the introduction of an inert gas will affect the total pressures *and* the partial pressures of each individual species.

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| **Example**:  A 1.0 L vessel is charged with 1.00 atm of A, and the following reaction is allowed to come to equilibrium at 298 K.  A(g) ↔ 2 B(g) Kp = 3.10   1. What are the equilibrium partial pressures and mole fractions of A and B? 2. If the volume of the container is doubled, what are the equilibrium partial pressures and mole fractions of A and B? 3. If 1.000 atm of Ar (an inert gas) is introduced into the system described in b), what are the equilibrium partial pressures and mole fractions of A and B once equilibrium is reestablished?   **Solution**:  First, we can use an ICE[[1]](#footnote-1) table to solve part a).   |  |  |  | | --- | --- | --- | |  | A | 2 B | | Initial | 1.00 atm | 0 | | Change | -x | +2x | | Equilibrium | 1.00 atm - x | 2x |   So (for convenience, consider Kp to have units of atm)  Solving for x yields values of  Clearly, x1, while a solution to the mathematical problem, is not physically meaningful since the equilibrium pressure of B cannot be negative. So the equilibrium partial pressures are given by  So the mole fractions are given by  In part b) the volume is doubled. Again, an ICE table is useful. The initial pressures will be half of the equilibrium pressures found in part a).   |  |  |  | | --- | --- | --- | |  | A | 2 B | | Initial | 0.213 atm | 0.574 atm | | Change | -x | +2x | | Equilibrium | 0.213 atm - x | 0.574 atm + 2x |   So the new equilibrium pressures can be found from  And the values of x that solve the problem are  We reject the negative root (since it would cause both of the partial pressures to become negative. So the new equilibrium partial pressures are  And the mole fractions are  We can see that the mole fraction of A decreased and the mole fraction B increased. This is the result expected by Le Chatlier’s principle since the lower total pressure favors the side of the reaction with more moles of gas.  Now for part c). We introduce 1.000 atm of an inter gas. The new partial pressures are  pA = 0.154 atm, pB = 0.692 atm, pAr = 1.000 atm  And because the partial pressures of A and B are unaffected, the equilibrium does not shift! What *is* affected is the composition, and so the mole fractions will change.  And since  Within round-off error, the value obtained is the equilibrium constant. So the conclusion is that the introduction of an inert gas, even though it increases the total pressure, does not induce a change in the partial pressures of the reactants and products, so it does not cause the equilibrium to shift. |

### Degree of Dissociation

Reactions such as the one in the previous example involve the dissociation of a molecule. Such reactions can be easily described in terms of the fraction of reactant molecules that actually dissociate to achieve equilibrium in a sample. This fraction is called the **degree of dissociation**. For the reaction in the previous example

A(g) ↔ 2 B(g)

the degree of dissociation can be used to fill out an ICE table. If the reaction is started with n moles of A, and  is the fraction of A molecules that dissociate, the ICE table will look as follows.

|  |  |  |
| --- | --- | --- |
|  | A | 2 B |
| Initial | n | 0 |
| Change | -n | +2n |
| Equilibrium | n(1 - ) | 2n |

The mole fractions of A and B can then be expressed by

Based on these mole fractions

And so Kp, which can be expressed as

is given by

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| **Example**:  Based on the values given below, find the equilibrium constant at 25 oC and degree of dissociation for a system that is at a total pressure of 1.00 atm for the reaction  N2O4(g) ↔ 2 NO2(g)   |  |  |  | | --- | --- | --- | |  | N2O4(g) | NO2(g) | | (kJ/mol) | 99.8 | 51.3 |   First, the value of Kp can be determined from .  So, using ,  The degree of dissociation can then be calculated from  Solving for ,  *Note*: since  represents the fraction of N2O4 molecules dissociated, it must be a positive number between 0 and 1. |

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| **Example**:  Consider the gas-phase reaction  A + 2B ↔ 2C  A reaction vessel is initially filled with 1.00 mol of A and 2.00 mol of B. At equilibrium, the vessel contains 0.60 mol C and a total pressure of 0.890 atm at 1350 K.   1. How many mol of A and B are present at equilibrium? 2. What is the mole fraction of A, B, and C at equilibrium? 3. Find values for Kx, Kp, and   **Solution**:  Let’s build an ICE table!   |  |  |  |  | | --- | --- | --- | --- | |  | A | 2 B | 2 C | | Initial | 1.00 mol | 2.00 mol | 0 | | Change | -x | -2x | +2x | | Equilibrium | 1.00 mol - x | 2.00 mol – 2x | 2x = 0.60 mol |   From the equilibrium measurement of the number of moles of C, x = 0.30 mol. So at equilibrium,   |  |  |  |  | | --- | --- | --- | --- | |  | A | 2 B | 2 C | | Equilibrium | 0.70 mol | 1.40 mol | 0.60 mol |   The total number of moles at equilibrium is 2.70 mol. From these data, the mole fractions can be determined.  So Kx is given by  And Kp is given by , so  The thermodynamic equilibrium constant is unitless, pf course, since the pressures are all divided by 1 atm. So the actual value of Kp is 0.794. This value can be used to calculate using  So |

## Temperature Dependence of Kp

The value of Kp is independent of pressure, although the composition of a system at equilibrium may be very much dependent on pressure. Temperature dependence is another matter. Because the value of is dependent on temperature, the value of Kp is as well. The form of the temperature dependence can be taken from the definition of the Gibbs function. At constant temperature and pressure

Substituting

For the two values of Go and using the appropriate temperatures, yields

And simplifying the expression so that only terms involving K are on the left and all other terms are on the right results in the **van’t Hoff equation**, which describes the temperature dependence of the equilibrium constant.

Because of the assumptions made in the derivation of the Gibbs-Helmholtz equation, this relationship only holds if Ho is independe of temperature over the range being considered. This expression also suggests that a plot of ln(K) as a function of 1/T should produce a straight line with a slope equal to –Ho/R. Such a plot is known as a **van’t Hoff plot**, and can be used to determine the reaction enthalpy.

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| **Example**:  A certain reaction has a value of Kp = 0.0260 at 25 oC and . Calculate the value of Kp at 37 oC.  **Solution**:  This is a job for the van’t Hoff equation!  T1 = 298 K K1 = 0.0260 =32400 J/mol  T2 = 310 K K2 = ?  So the equation becomes  *Note*: the value of K2 increased with increasing temperature, which is what is expected for an endothermic reaction. An increase in temperature should result in an increase of product formation in the equilibrium mixture. But unlike a change in pressure, a change in temperature actually leads to a change in the value of the equilibrium constant! |

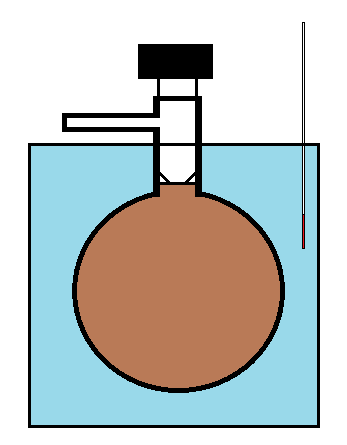
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| **Example**:  Given the following average bond enthalpies for P-Cl and Cl-Cl bonds, predict whether or not an increase in temperature will lead to a larger or smaller degree of dissociation for the reaction  PCl5 ↔ PCl3 + Cl2   |  |  | | --- | --- | | X-Y | D(X-Y) (kJ/mol) | | P-Cl | 326 | | Cl-Cl | 240 |   **Solution**:  The estimated reaction enthalpy is given by the total energy expended breaking bonds minus the energy recovered by the formation of bonds. Since this reaction involves breaking two P-Cl bonds (costing 652 kJ/mol) and the formation of one Cl-Cl bon (recovering 240 kJ/mol), it is clear that the reaction is endothermic (by approximately 412 kJ/mol). As such, an increase in temperature should increase the value of the equilibrium constant, causing the degree of dissociation to be increased at the higher temperature. |

### The Dumas Bulb Method for Measuring Decomposition Equilibrium

A classic example of an experiment that is employed in many physical chemistry laboratory courses uses a Dumas Bulb method to measure the dissociation of N2O4(g) as a function of temperature (Mack & France, 1934). In this experiment, a glass bulb is used to create a constant volume container in which a volatile substance can evaporate, or achieve equilibrium with other gases present. The latter is of interest in the case of the reaction

N2O4(g) ↔ 2 NO2(g)

The reaction is endo thermic, so at higher temperatures, a larger degree of dissociation is observed.



The procedure is to first calibrate the internal volume of the Dumas bulb. This is done using a heavy gas (such as SF6) and comparing the mass of the bulb when evacuated to the mass of the bulb full of the calibrant gas at a particular pressure and temperature.

The Dumas bulb is then charged with a pure sample of the gas to be investigated (such as N2O4) and placed in a thermalized bath. It is then allowed to come to equilibrium. Once Equilibrium is established, the stopcock is opened to allow gas to escape until the internal pressure is set to the pressure of the room. The stopcock is then closed and the bulb weighed to determine the total mass of gas remaining inside. The experiment is repeated at higher and higher temperatures (so that at each subsequent measurement, the larger degree of dissociation creates more molecules of gas and an increase in pressure in the bulb (along with the higher temperature), which then leads to the expulsion of gas when the pressure is equilibrated.

The degree of dissociation is then determined based on the calculated gas density at each temperature.

where 1 is the measured density and 2 is the theoretical density if no dissociation occurs (calculated from the ideal gas law for the given temperature, pressure, and molar mass of the dissociating gas) and n is the number of fragments into which the dissociating gas dissociates (in the case of N2O4 ↔ 2 NO2, n = 2.) The equilibrium constant is then calculated as

Finally, a van’t Hoff plot is generated to determine the reaction enthalpy.

## Acid-Base Equilibria

A great many processes involve proton transfer, or acid-base types of reactions. As many biological systems depend on carefully controlled pH, these types of processes are extremely important. The pH is defined by

where [H+] is in mol/L.

The dissociation of a weak acid in water is governed by the equilibrium defined by

HA(aq) ↔ H+(aq) + A-(aq)

The equilibrium constant for such a reaction, Ka, takes the form

As is the case for all thermodynamic equilibrium constants, the concentrations are replaced by activities and the equilibrium constant is unitless. However, if all species behave ideally (have unit activity coefficients) the units can be used as a very useful guide in solving problems.

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| **Example**:  What is the pH of a 0.200 M HOAc (acetic acid) solution? (Ka = 1.8 x 10-5 M)  **Solution**:  An ICE table will come in very handy here!   |  |  |  |  | | --- | --- | --- | --- | |  | HOAc | H+ | OAc- | | Initial | 0.200 M | 0 | 0 | | Change | -x | +x | +x | | Equilibrium | 0.200 M - x | x | x |     The equilibrium problem can then be set up as  Substituting the values that are known  This produces a quadratic equation, and thus two values of x which satisfy the relationship.  The negative root is not physically meaningful since the concentrations of H+ and OAc- cannot be negative. Using the value of x2 as [H+], the pH is calculated to be |

### The Auto-ionization of Water

Water is a very important solvent as water molecules have large dipole moments which create favorable interactions with ionic compounds. Water also has a large dielectric constant which damps the electric field generated by ions in solutions, making the comparative interactions with water more favorable than with other ions in solution in many cases. But water also dissociates into ions through the reaction

H2O(l) ↔ H+(aq) + OH-(aq)

The equilibrium constant governing this dissociation is highly temperature dependent. The data below are presented by Bandura and Lvov (Bandura & Lvov, 2006)

|  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- |
| T (oC) | 0 | 25 | 50 | 75 | 100 |
| pKw | 14.95 | 13.99 | 13.26 | 12.70 | 12.25 |

From these data, a van’t Hoff plot can be constructed.

There is some curvature to the line, suggesting some (albeit small) temperature dependence for Hrxn. However, from the fit of these data, a value of Hrxn can be determined to be 52.7 kJ/mol. Of particular note is that the dissociation is endothermic, so increases in temperature will lead to a greater degree of dissociation.

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| **Example**:  What is the pH of neutral water at 37 oC (normal human body temperature)?  **Solution**:  From the best-fit line in the above van’t Hoff plot, the value of Kw can be calculated:  Since Kw gives the product of [H+] and [OH-] (which must be equal in a neutral solution),  And the pH is given by  *Note*: This is slightly less than a pH of 7.00, which is normally considered to be “neutral.” But a pH of 7.00 is only neutral at 25 oC! At higher temperatures, neutral pH is a lower value due to the endothermic nature of the auto-ionization water. |

### The Hydrolysis of a Weak Base

Hydrolysis is defined as a reaction with water that splits a water molecule. The hydrolysis of a weak base defines the equilibrium constant Kb.

A- + H2O ↔ HA + OH-‑

For this reaction, the equilibrium constant Kb is given by

The concentration (or activity) of the pure compound H2O is not included in the equilibrium expression because, being a pure compound in its standard state, it has unit activity throughout the process of establishing equilibrium. Further, it should be noted that when Kb is combined with the expression for Ka for the weak acid HA,

As a consequence, if one knows Ka for a weak acid, one also knows Kb for its conjugate base, since the product results in Kw.

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| **Example**:  What is the pH of a 0.150 M solution of KF? (For HF, pKa = 3.17 at 25 oC)  **Solution**:  The problem involves the hydrolysis of the conjugate base of HF, F-. The hydrolysis reaction is  F- + H2O ↔ HF + OH-  An ICE table is in order here.   |  |  |  |  | | --- | --- | --- | --- | |  | F- | HF | OH- | | Initial | 0.150 M | 0 | 0 | | Change | -x | +x | +x | | Equilibrium | 0.150 M - x | x | x |   So the expression for Kb is  In this case, the small value of Kb insures that the value of x will be negligibly small compared to 0.150 M. In this limit, the value of x (which is equal to [OH­-])  So [H+] is given by  And the pH is given by  *Note*: The pH of this salt solution is slightly basic. This is to be expected as KF can be thought of being formed in the reaction of a weak acid (HF) with a strong base (KOH). In the competition to control the pH, the strong base ends up winning the battle. |

### Buffers

Buffer solutions, which are of enormous importance in controlling pH in various processes, can be understood in terms of acid/base equilibrium. A buffer is created in a solution which contains both a weak acid and its conjugate base. This creates to absorb excess H+ or supply H+ to replace what is lost due to neutralization. The calculation of the pH of a buffer is straightforward using an ICE table approach.

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| Example:  What is the pH of a solution that is 0.150 M in KF and 0.250 M in HF?  Solution:  The reaction of interest is  HF ↔ H+ + F-  Let’s use an ICE table!   |  |  |  |  | | --- | --- | --- | --- | |  | HF | H+ | F- | | Initial | 0.250 M | 0 | 0.150 M | | Change | -x | +x | +x | | Equilibrium | 0.250 M - x | x | 0.150 M + x |   This expression results in a quadratic relationship, leading to two values of x that will make it true. Rejecting the negative root, the remaining root of the equation indicates  So the pH is given by |

For buffers made from acids with sufficiently large values of pKa the buffer problem can be simplified since the concentration of the acid and its conjugate base will be determined by their pre-equilibrium values. In these cases, the pH can be calculated using the Henderson-Hasselbalch equation.

If one considers the expression for Ka

Taking the log of both sides and multiplying by -1 yields

An rearrangement produces the form of the Henderson-Hasselbalch equation.

It should be noted that this expression will fail if 1) the pKa is too small and/or the concentrations [A-] or [HA] are too small, since the equilibrium concentration will deviate wildly from the pre-equilibrium values under these conditions.

## Solubility of Ionic Compounds

The solubility of ionic compounds in water can also be described using the concepts of equilibrium. If you consider the dissociation of a generic salt MX

MX(s) ↔ M+(aq) + X-(aq)

The equilibrium expression is

Ksp is the **solubility product** and is the equilibrium constant that describes the solubility of an electrolyte. And again, the pure solid MX is not included in the expression since it has unit activity throughout the establishment of equilibrium.

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| **Example**:  What is the maximum solubility of CuS at 25 oC? (Ksp = 1∙10-36 M2)  **Solution**:  Yup – time for an ICE table.   |  |  |  |  | | --- | --- | --- | --- | |  | CuS | Cu2+ | S2- | | Initial |  | 0 | 0 | | Change |  | +x | +x | | Equilibrium |  | x | x |   So the equilibrium expression is |

|  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |
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| **Example**:  What is the maximum solubility of CuS at 25 oC in 0.100 M NaS? (Ksp = 1∙10-36 M2)  **Solution**:  In this problem we need to consider the existence of S2-(aq) from the complete dissociation of the strong electrolyte NaS. An ICE table will help, as usual.   |  |  |  |  | | --- | --- | --- | --- | |  | CuS | Cu2+ | S2- | | Initial |  | 0 | 0.100 M | | Change |  | +x | +x | | Equilibrium |  | x | 0.100 M + x |   Given the miniscule magnitude of the solubility product, x will be neglibly small compared to 0.100 MSo the equilibrium expression is |

The huge reduction in solubility is due to the **common ion effect**. The existence of sulfide in the solution due to sodium sulfide greatly reduces the solutions capacity to support additional sulfide due to the dissociation of CuS.

## Relating Keq to Ho and So

At constant temperature and pressure

Noting that

it is evident that

Dividing both sides by –RT produces

This result suggests that a plot of ln(K) as a function of 1/T should produce a line with a slope equal to –Ho/R and an intercept of So/R. Such a plot using Ksp for Ca(OH)2 is shown below.

The result shows that the dissociation is exothermic.

It also shows a negative change in entropy upon dissociation, which is counter intuitive.

The explanation for the negative entropy change is that for certain ions in solution that the organization of water molecules within the solvation sphere leads to a large negative entropy change for the solvent – particularly at the saturation point. The negative entropy change for the solvent, at least in this case, overcomes the positive entropy change for the solute, leading to a net negative entropy change.

Another approach to determining the entropy change is to recognize

So a plot of vs. T should yield a curve, where the slope at any point gives . Below is a plot of the data calculated from Ksp for Ca(OH)2 as a function of temperature. The data are fairly linear, meaning that is relatively independent of temperature over this range. The plot also confirms the negative entropy change for the reaction

Ca(OH)2 ↔ Ca2+(aq) + 2 OH-(aq)

The magnitude of the entropy change determined in this manner is very similar to that determined from the ln(Ksp) vs 1/T method, as expected for a system in which is independent of temperature.

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# Learning Objectives

After mastering the material covered in this chapter, one will be able to:

1. Know stuff.
2. Do stuff.
3. Be stuff.

# Problems

Oh – I got problems . . .

1. ICE is an acronym for “Initial, Change, Equilibrium”. An ICE table is a tool that is used to solve equilibrium problems in terms of an unknown number of moles (or something proportional to moles, such as pressure or concentration) will shift for a system to establish equilibrium. See (Tro, 2014) or a similar General Chemistry text for more background and information. [↑](#footnote-ref-1)