

## 8.1: Solutions of Acids or Bases Containing a Common Ion

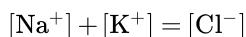
### Learning Objectives

- Recognize common ions from various salts, acids, and bases.
- Calculate concentrations involving common ions.
- Calculate ion concentrations involving chemical equilibrium.

The common-ion effect is used to describe the effect on an equilibrium when one or more species in the reaction is shared with another reaction. This results in a shifting of the equilibrium properties.

### Introduction

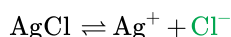
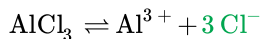
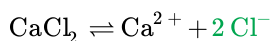
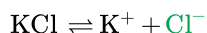
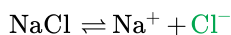
The solubility products  $K_{sp}$ 's are equilibrium constants in heterogeneous equilibria (i.e., between two different phases). If several salts are present in a system, they all ionize in the solution. If the salts contain a common cation or anion, these salts contribute to the concentration of the common ion. Contributions from all salts must be included in the calculation of concentration of the common ion. For example, a solution containing sodium chloride and potassium chloride will have the following relationship:



Consideration of *charge balance* or *mass balance* or both leads to the same conclusion.

### Common Ions

When NaCl and KCl are dissolved in the same solution, the  $\text{Cl}^-$  ions are *common* to both salts. In a system containing NaCl and KCl, the  $\text{Cl}^-$  ions are common ions.



For example, when AgCl is dissolved into a solution already containing NaCl (actually  $\text{Na}^+$  and  $\text{Cl}^-$  ions), the  $\text{Cl}^-$  ions come from the ionization of both AgCl and NaCl. Thus,  $[\text{Cl}^-]$  differs from  $[\text{Ag}^+]$ . The following examples show how the concentration of the common ion is calculated.

### ✓ Example 8.1.1

What are  $[\text{Na}^+]$ ,  $[\text{Cl}^-]$ ,  $[\text{Ca}^{2+}]$ , and  $[\text{H}^+]$  in a solution containing 0.10 M each of NaCl,  $\text{CaCl}_2$ , and HCl?

#### Solution

Due to the conservation of ions, we have

$$[\text{Na}^+] = [\text{Ca}^{2+}] = [\text{H}^+] = 0.10 \text{ M.}$$

but

$$\begin{aligned} [\text{Cl}^-] &= 0.10 \text{ (due to NaCl)} \\ &+ 0.20 \text{ (due to CaCl}_2\text{)} \\ &+ 0.10 \text{ (due to HCl)} \\ &= 0.40 \text{ M} \end{aligned}$$

### ? Exercise 8.1.1

John poured 10.0 mL of 0.10 M NaCl, 10.0 mL of 0.10 M KOH, and 5.0 mL of 0.20 M HCl solutions together and then he made the total volume to be 100.0 mL. What is  $[Cl^-]$  in the final solution?

$$[Cl^-] = \frac{0.1 \text{ M} \times 10 \text{ mL} + 0.2 \text{ M} \times 5.0 \text{ mL}}{100.0 \text{ mL}} = 0.020 \text{ M}$$

Le Chatelier's Principle states that if an equilibrium becomes unbalanced, the reaction will shift to restore the balance. If a common ion is added to a weak acid or weak base equilibrium, then the equilibrium will shift towards the reactants, in this case the weak acid or base.

### ✓ Example 8.1.2

Consider the lead(II) ion concentration in this *saturated* solution of  $PbCl_2$ . The balanced reaction is



Defining  $s$  as the concentration of dissolved lead(II) chloride, then:

$$[Pb^{2+}] = s$$

$$[Cl^-] = 2s$$

These values can be substituted into the solubility product expression, which can be solved for  $s$ :

$$\begin{aligned} K_{sp} &= [Pb^{2+}][Cl^-]^2 \\ &= s \times (2s)^2 \\ 1.7 \times 10^{-5} &= 4s^3 \\ s^3 &= \frac{1.7 \times 10^{-5}}{4} \\ &= 4.25 \times 10^{-6} \\ s &= \sqrt[3]{4.25 \times 10^{-6}} \\ &= 1.62 \times 10^{-2} \text{ mol dm}^{-3} \end{aligned}$$

The concentration of lead(II) ions in the solution is  $1.62 \times 10^{-2} \text{ M}$ . Consider what happens if sodium chloride is added to this saturated solution. Sodium chloride shares an ion with lead(II) chloride. The chloride ion is **common** to both of them; this is the origin of the term "common ion effect".

Look at the original equilibrium expression in Equation 8.1.1. What happens to that equilibrium if extra chloride ions are added? According to *Le Chatelier*, the position of equilibrium will shift to counter the change, in this case, by removing the chloride ions by making extra solid lead(II) chloride.

Of course, the concentration of lead(II) ions in the solution is so small that only a tiny proportion of the extra chloride ions can be converted into solid lead(II) chloride. The lead(II) chloride becomes even *less soluble*, and the concentration of lead(II) ions in the solution *decreases*. This type of response occurs with any sparingly soluble substance: it is less soluble in a solution which contains any ion which it has in common. This is the common ion effect.

### A Simple Example

If an attempt is made to dissolve some lead(II) chloride in some 0.100 M sodium chloride solution instead of in water, what is the equilibrium concentration of the lead(II) ions this time? As before, define  $s$  to be the concentration of the lead(II) ions.

$$[Pb^{2+}] = s$$

The calculations are different from before. This time the concentration of the chloride ions is governed by the concentration of the sodium chloride solution. The number of ions coming from the lead(II) chloride is going to be tiny compared with the 0.100 M

coming from the sodium chloride solution.

In calculations like this, it can be assumed that the concentration of the common ion is entirely due to the other solution. This simplifies the calculation.

So we assume:

$$[\text{Cl}^-] = 0.100 \text{ M}$$

The rest of the mathematics looks like this:

$$\begin{aligned} K_{sp} &= [\text{Pb}^{2+}][\text{Cl}^-]^2 \\ &= s \times (0.100)^2 \\ 1.7 \times 10^{-5} &= s \times 0.0100 \end{aligned}$$

therefore:

$$\begin{aligned} s &= \frac{1.7 \times 10^{-5}}{0.0100} \\ &= 1.7 \times 10^{-3} \text{ M} \end{aligned}$$

Finally, compare that value with the simple saturated solution:

Original solution:

$$[\text{Pb}^{2+}] = 0.0162 \text{ M}$$

Solution in 0.100 M NaCl solution:

$$[\text{Pb}^{2+}] = 0.0017 \text{ M}$$

The concentration of the lead(II) ions has decreased by a factor of about 10. If more concentrated solutions of sodium chloride are used, the solubility decreases further.

Adding a common ion to a system at equilibrium affects the equilibrium composition, but **not** the ionization constant.

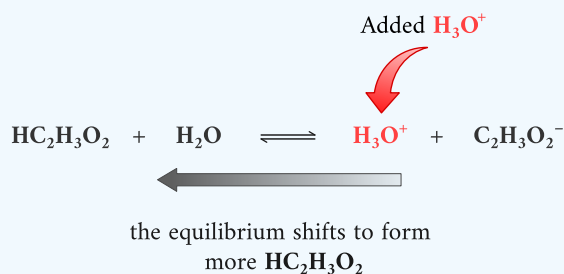
## Common Ion Effect with Weak Acids and Bases

Adding a common ion prevents the weak acid or weak base from ionizing as much as it would without the added common ion. The common ion effect suppresses the ionization of a weak acid by adding more of an ion that is a product of this equilibrium.

### ✓ Example 8.1.3A

The common ion effect of  $\text{H}_3\text{O}^+$  on the ionization of acetic acid

When a strong acid supplies the common ion  $\text{H}_3\text{O}^+$  the equilibrium shifts to form more  $\text{HC}_2\text{H}_3\text{O}_2$ .



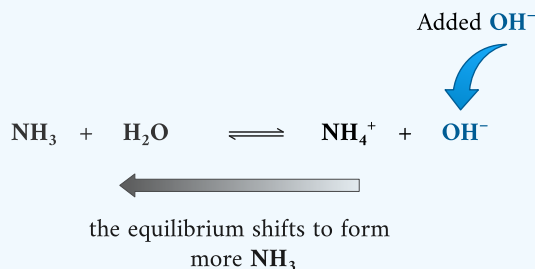
The common ion effect can be conceptualized as an applied Le Chatelier effect. Adding products will shift the equilibrium to more favor the reactants. (CC BY-NC 4.0; Ümit Kaya via LibreTexts)

The common ion effect suppresses the ionization of a weak base by adding more of an ion that is a product of this equilibrium.

### ✓ Example 8.1.3B

Consider the common ion effect of  $\text{OH}^-$  on the ionization of ammonia

When a strong base supplies the common ion  $\text{OH}^-$  the equilibrium shifts to form more  $\text{NH}_3$ .



The common ion effect can be conceptualized as an applied Le Chatelier effect. Adding products will shift the equilibrium to more favor the reactants. (CC BY-NC 4.0; Ümit Kaya via LibreTexts)

Adding the common ion of hydroxide shifts the reaction towards the left to decrease the stress (in accordance with Le Chatelier's Principle), forming more reactants. This decreases the reaction quotient, because the reaction is being pushed towards the left to reach equilibrium. The equilibrium constant,  $K_b = 1.8 \times 10^{-5}$ , does not change. The reaction is put out of balance, or equilibrium.

$$Q_a = \frac{[\text{NH}_4^+][\text{OH}^-]}{[\text{NH}_3]}$$

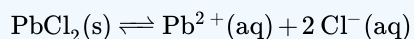
At first, when more hydroxide is added, the quotient is greater than the equilibrium constant. The reaction then shifts right, causing the denominator to increase, decreasing the reaction quotient and pulling towards equilibrium and causing  $Q$  to decrease towards  $K$ .

## Common Ion Effect on Solubility

Adding a common ion decreases solubility, as the reaction shifts toward the left to relieve the stress of the excess product. Adding a common ion to a dissociation reaction causes the equilibrium to shift left, toward the reactants, causing precipitation.

### ✓ Example 8.1.4

Consider the reaction:



What happens to the solubility of  $\text{PbCl}_2(\text{s})$  when 0.1 M NaCl is added?

**Solution**

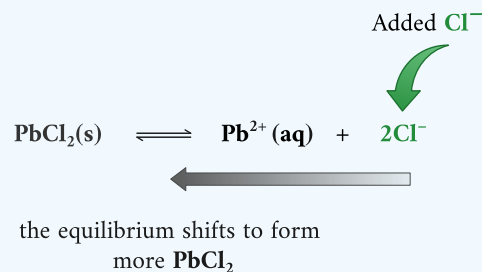
$$K_{sp} = 1.7 \times 10^{-5}$$

$$Q_{sp} = 1.8 \times 10^{-5}$$

Identify the common ion:  $\text{Cl}^-$

Notice:  $Q_{sp} > K_{sp}$  The addition of NaCl has caused the reaction to shift out of equilibrium because there are more dissociated ions. Typically, solving for the molarities requires the assumption that the solubility of  $\text{PbCl}_2(\text{s})$  is equivalent to the concentration of  $\text{Pb}^{2+}$  produced because they are in a 1:1 ratio.

When a salt supplies the common ion  $\text{Cl}^-$  the equilibrium shifts to form more  $\text{PbCl}_2$ .



The common ion effect can be conceptualized as an applied Le Chatelier effect. Adding products will shift the equilibrium to more favor the reactants. (CC BY-NC 4.0; Ümit Kaya via LibreTexts)

Because  $K_{sp}$  for the reaction is  $1.7 \times 10^{-5}$ , the overall reaction would be

$$(s)(2s)^2 = 1.7 \times 10^{-5}.$$

Solving the equation for  $s$  gives  $s = 1.62 \times 10^{-2} \text{ M}$ . The coefficient on  $\text{Cl}^-$  is 2, so it is assumed that twice as much  $\text{Cl}^-$  is produced as  $\text{Pb}^{2+}$ , hence the '2s.' The solubility equilibrium constant can be used to solve for the molarities of the ions at equilibrium.

The molarity of  $\text{Cl}^-$  added would be 0.1 M because  $\text{Na}^+$  and  $\text{Cl}^-$  are in a 1:1 ratio in the ionic salt, NaCl. Therefore, the overall molarity of  $\text{Cl}^-$  would be  $2s + 0.1$ , with  $2s$  referring to the contribution of the chloride ion from the dissociation of lead chloride.

$$\begin{aligned} Q_{sp} &= [\text{Pb}^{2+}][\text{Cl}^-]^2 \\ &= 1.8 \times 10^{-5} \\ &= (s)(2s + 0.1)^2 \\ s &= [\text{Pb}^{2+}] \\ &= 1.8 \times 10^{-3} \text{ M} \\ 2s &= [\text{Cl}^-] \\ &\approx 0.1 \text{ M} \end{aligned}$$

Notice that the molarity of  $\text{Pb}^{2+}$  is lower when NaCl is added. The equilibrium constant remains the same because of the increased concentration of the chloride ion. To simplify the reaction, it can be assumed that  $[\text{Cl}^-]$  is approximately 0.1 M since the formation of the chloride ion from the dissociation of lead chloride is so small. The reaction quotient for  $\text{PbCl}_2(\text{s})$  is greater than the equilibrium constant because of the added  $\text{Cl}^-$ . This therefore shift the reaction left towards equilibrium, causing precipitation and lowering the current solubility of the reaction. Overall, the solubility of the reaction decreases with the added sodium chloride.

## References

1. Harwood, William S., F. G. Herring, Jeffery D. Madura, and Ralph H. Petrucci. General Chemistry Principles and Modern Applications. 9th ed. New Jersey: Prentice Hall, 2007.

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