learning objectives

• Define hydrates and identify common hydrates
• Outline properties of hydrates
• Determine formulas for hydrates

Introduction

Water, the most common chemical on earth, can be found in the atmosphere as water vapor. Some chemicals, when exposed to water in the atmosphere, will reversibly either adsorb it onto their surface or include it in their structure forming a complex in which water generally bonds with the cation in ionic substances. The water present in the latter case is called water of hydration or water of crystallization. Common examples of minerals that exist as hydrates are gypsum (\(\text{CaSO}_4 \cdot 2\text{H}_2\text{O}\)), Borax (\(\text{Na}_3\text{B}_4\text{O}_7 \cdot 10\text{H}_2\text{O}\)) and Epsom salts (\(\text{MgSO}_4 \cdot 7\text{H}_2\text{O}\)). Hydrates generally contain water in stoichiometric amounts; hydrates’ formulae are represented using the formula of the anhydrous (non-water) component of the complex followed by a dot then the water (\(\text{H}_2\text{O}\)) preceded by a number corresponding to the ratio of \(\text{H}_2\text{O}\) moles per mole of the anhydrous component present. They are typically named by stating the name of the anhydrous component followed by the Greek prefix specifying the number of moles of water present then the word hydrate (example: \(\text{MgSO}_4 \cdot 7\text{H}_2\text{O}\): magnesium sulfate heptahydrate).

Properties of Hydrates

It is generally possible to remove the water of hydration by heating the hydrate.

The residue obtained after heating, called the anhydrous compound, will have a different structure and texture and may have a different color than the hydrate.

Example:

\[
\underbrace{\text{CuSO}_4 \cdot 5\text{H}_2\text{O} (s)}_{\text{Deep Blue}} \xrightarrow{\Delta} \underbrace{\text{CuSO}_4 (s)}_{\text{Ashy White}} + 5\text{H}_2\text{O} (g) \label{1}
\]

\[
\underbrace{\text{CuSO}_4 (s)}_{\text{Ashy White}} \xrightarrow{\ce{H_2O (l)}} \underbrace{\text{CuSO}_4 (aq)}_{\text{Deep Blue}} \label{2}
\]

Any anhydrous compound from a hydrate generally has the following properties:

• Highly soluble in water
• When dissolved in water, the anhydrous compound will have a color similar to that of the original hydrate even if it had changed color going from the hydrate to the anhydrous compound.

Most hydrates are stable at room temperature. However, some spontaneously lose water upon standing in the atmosphere, they are said to be **efflorescent**.

Other compounds can spontaneously absorb water from the surrounding atmosphere, they are said to be **hygroscopic**. Some hygroscopic substances, such as \(\ce{P2O5}\) and anhydrous \(\ce{CaCl2}\), are widely used to "dry" liquids and gases (see experiment on the Molecular Weight of \(\ce{CO2}\)); they are referred to as desiccants. Other hygroscopic substances, such as solid \(\ce{NaOH}\), absorb so much water from the atmosphere that they dissolve in this water, these substances are said to be **deliquescent**. Some compounds like carbohydrates release water upon heating by decomposition of the compound rather than by loss of the water of hydration. These compounds are not considered true hydrates as the hydration process is not reversible.

### Formula of a Hydrate \((\text{Anhydrous Solid}\ce{*}x\ce{H2O})\)

The formula of a hydrate can be determined by dehydrating a known mass of the hydrate, then comparing the masses of the original hydrate and the resulting anhydrous solid. The mass of water evaporated is obtained by subtracting the mass of the anhydrous solid from the mass of the original hydrate (Equation 3):

\[
m_{\ce{H2O}} = m_{\text{Hydrate}} - m_{\text{Anhydrous Solid}} \label{3}
\]

From the masses of the water and anhydrous solid and the molar mass of the anhydrous solid (the formula of the anhydrous solid will be provided), the number of moles of water and moles of the anhydrous solid are calculated as shown below (Equations 4, 5):

\[
n_{\ce{H2O}} = \frac{m_{\ce{H2O}}}{MM_{\ce{H2O}}} \label{4}
\]

\[
n_{\text{Anhydrous Solid}} = \frac{m_{\text{Anhydrous Solid}}}{MM_{\text{Anhydrous Solid}}} \label{5}
\]

In order to determine the formula of the hydrate, \((\text{Anhydrous Solid}\ce{*}x\ce{H2O})\), the number of moles of water per mole of anhydrous solid \((x)\) will be calculated by dividing the number of moles of water by the number of moles of the anhydrous solid (Equation 6).

\[
x = \frac{n_{\ce{H2O}}}{n_{\text{Anhydrous Solid}}} \label{6}
\]
Exercise \((\PageIndex{1})\)

You weight out a 0.470-sample of hydrated nickel(II) chloride, \(\text{NiCl}_2 \cdot x\text{H}_2\text{O}\). Upon heating, the mass of the anhydrous salt that remains is 0.256 grams. What is the formula of the hydrate? What is the name of the hydrate?

**Answer**

\[
n \text{(moles of H}_2\text{O)} = \frac{0.256 \text{ g H}_2\text{O}}{18.0152 \text{ g H}_2\text{O}} = 0.013876 \text{ moles of H}_2\text{O}
\]

\[
n \text{(moles of NiCl}_2) = \frac{0.470 \text{ g NiCl}_2}{139.743 \text{ g H}_2\text{O}} = 0.0029753 \text{ moles of NiCl}_2
\]

\[
n \text{(moles of NiCl}_2) = \frac{0.013876 \text{ moles of H}_2\text{O}}{6.0} = 0.00010740 \text{ moles of NiCl}_2
\]

So, \(\text{NiCl}_2 \cdot 6\text{H}_2\text{O}\)

Name: Nickel(II) chloride hexahydrate

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**Contributors:**

- Original Article can be found here: [S: Properties of Hydrates (Experiment)]
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