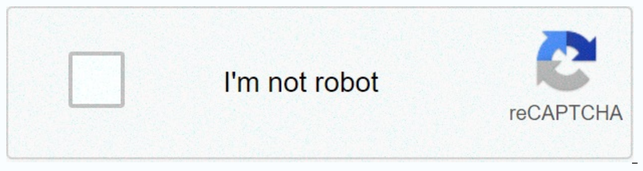

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Formula of a hydrate lab answers

Background: Forms of Matter Ionic Hydrates Procedure Introduction to quantitative analysis: determination of the proportion of subunits in a chemical sample Laboratory tool: Bunsen Burner Introduction to forms of matter with the example of ionic hydrates Elements: Matter consists of only one type of atom (e.g. Na2SO4 (ion compound; Na2SO4 is a unit of formula) H2O (covalent; exists as separate H2O molecules) Molecules: Matter consists of more than one atom; atoms are connected by covalent bonds (e.g. Ionic compounds (salts) in which one or more water molecules are bound in the crystalline salt structure: CoCl2.6H2O (anhydrous salt)(hydration water) BaCl2.2H2O CaCO3.2 1/2 H2O CaSO4.5H2O There is a unique ratio of water molecules to formula units in the crystalline structure. Interactions in ion hydrate bacl2.3H2O Procedure (Review) Procedure (Details) Each student will work individually. Wipe the crucible thoroughly and cover with a clean cloth towel to remove dirt and other solid particles. Then determine the weight of the crucible (and cover) at +/- 0,01 g. Obtain about 1 g of the hydrate sample and transfer the sample to the crucible. The samples are located on the reagent counter; the identification code must be stored. Tubes containing 1 g of hydrate samples are also located on the reagent counter. To quickly obtain about 1 g of hydrate sample, take the tube from the cabinet and obtain about the same amount of hydration from the reagent bottle as can be found in one of the reference samples. After sampling, quickly re-attach the cap to the reagent bottle and tighten it safely. Transfer the hydrate sample to the crucible. Determine the weight of the crucible, the moistened sample and the crucible cover. Using crucible pliers, place the crucible and contents back on the clay triangle. Partially cover the crucible opening with a lid. Heat the crucible and its contents with a small flame for 5 minutes. Increase the flame temperature and heat with medium flame for 5 min. Additionally, increase the flame and heat the sample for an additional 10 minutes. Do not allow the crucible to turn red. (Overheating can lead to sample decomposition!) Using crucible pliers, remove the crucible (with the lid in place) from the clay triangle and place on the wire gas on the laboratory bench. Allow the crucible to cool to room temperature. (Hold your hand about 1 cm above the crucible to test.) Then determine the weight of the crucible and the contents (and cover) at +/- 0,01 g. Heat the sample to a constant mass. Heat the crucible, with the lid tilted as before, for 10 min. with a hot flame. Allow the crucible to cool to room temperature and weigh again. Difference between crucible weights plus sample, after first and second should not be more than 0.03 g. If the difference is greater than 0,03 g, repeat the heating and cooling procedure until the difference between the successive heatings is less than this limit. Then you heated the sample to a constant mass. After heating the crucible and the contents to a constant mass, transfer the contents of the crucible to the waste container. You may want, if time allows, to make a second determination. The determination of the hydrate sample formula 1, 3 and 5 are magnesium sulphate hydrates, MgSO4.xH2O 2 and 4 samples are zinc sulphate hydrates, ZnSO4.xH2O To specify a formula, specify the following # mols H2O# mols H2O x = __ or y = __ # mols MgSO4# mols ZnSO4 To determine the number of salt anaesthest moles and H 2O: 1 mol MgSO4 # mols MgSO4 = g MgSO4 x ____ (measured mass anhydrous salt MgSO4)(1/molar) 1 mol H2O # mols H2O = g H2O x ___ (weight of lost water H2O) (1/molar mass) Return to the schedule of the laboratory chemical rules. Targets Identification of hydrates in the group of compounds Testing the properties of hydrates Determination of the number of water moths hydration in hydration Water, the most common chemical on earth, can be found in the atmosphere as water steam. Some chemicals exposed to water in the atmosphere will reversiblely adsorb it to its surface or incorporate it into its structure, creating a complex in which water usually binds to cation in ionic substances. The water present in the latter case is called hydration water or crystallization water. Typical examples of minerals that exist as hydrates are gypsum (\(\ce{CaSO4*2H2O}\)), Borax (\(\ce{Na3B4O7*10H2O}\)) and Epsom salts (\(\ce{MgSO4*7H2O}\)). Hydrates usually contain water in stoichiometric quantities; Hydrate patterns are represented by an anhydrous (non-ahydrous) formula of the complex component followed by a period followed by water (\(\ce{H2O}\)) preceded by a number corresponding to the mole proportion \(\ce{H2O}\) on the mole of the anhydrous component. They are usually named after the name of the anhydrous component followed by the Greek prefix for the number of water moths present, followed by the word hydrate (example: \(\ce{MgSO4*7H2O}\): magnesium sulfate heptahydrate). In general, it is possible to remove hydration water by heating the hydrate. The Le Chatelier principle stipulates that adding heat to an endothermic reaction (heat is a reactor) will shift the reaction to the right (product page). Heating will move the drainage equation below right because it is an endothermic reaction. The residue obtained after heating, called anhydrous compound, will have a different structure and consistency and may have a different color than the hydrate. Example: \[(s)}} _{\text{Deep Blue}} \ce{->[\Delta]} \underbrace{ \ce{CuSO4 (s)}_{\text{Ashy White}} \ce{+ 5 H2O (g)} \label{1}\] \[\underbrace{\ce{CuSO4 (s)}} _{\text{Ashy White}} \ce{->[\ce{H2O (l)}]} \underbrace{ \ce{CuSO4 (aq)}_{\text{Deep Blue}} \label{2}\] Any anhydrous relationship with a hydrate usually has the following properties: Highly water soluble When dissolved in water, anhydrous compound will have a color similar to the original hydration , even if it has changed color from hydration to anhydrous compound. Most hydrates are stable at room temperature. However, some spontaneously lose standing water in the atmosphere, it is said to bloom. 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 molecular weight \(\ce{CO2}\)); they are referred to as drying agents. Other hygroscopic substances, such as solids \(\ce{NaOH}\), absorb so much water from the atmosphere that they dissolve in this water, they are said to be deceptive. Some compounds, such as carbohydrates, release water when heated by decomposition of the compound, rather than by the loss of hydration water. These compounds are not considered real hydrates because the irrigation process is not reversible. The hydrate formula can be determined by dehydration of the known hydrate mass and then by comparing the masses of the primary hydrate and the resulting anhydrous solid. The mass of evaporated water is obtained by taking the anhydrous mass away from the mass of the primary hydrate (\ref{3}): \[m_{\ce{H2O}} = m_{\text{14 Waterline}} - m_{\text {Anhydrous Solid}} \label{3}\] From the masses of water and anhydrous solids and molar mass of the anhydrous solid (a pattern of the anhydrous solid will be given), the number of water moths and anhydrous moles of the solid is calculated as shown below (\ref{4} , \ref{5}): \[n_{\ce{H2O}} = \frac{m_{\ce{H2O}}{MM_{\ce{H2O}}} \label{4}\] \[n_ MM_{\text{Anhydrous Solid}} = \frac{m_{\text{Anhydrous Solid}}}{MM_{\text{Anhydrous Solid}}} \label{5}\] To specify a hydrate formula, [\(\text{Anhydrous solid}\ce{*}x\ce{H2O}\)] , the number of water moths per mole of anhydrous solid (\(x\)) will be calculated by dividing the number of water moths by the number of moths anhydrous solid (Equation \ref{6}). \[x = \frac{n_{\ce{H2O}}{n_{\text{Anhydrous solid}}} \label{6}\] Hardware materials and solutions : Nitric acid (6M) Solids: nickel (II) chloride, cobalt chloride (II), sacrose, calcium carbonate, barium chloride, sodium tetraborate, potassium chloride, sodium sulphate hydrate, iron chloride (III), potassium aluminium sulphate, calcium chloride, copper Unknown. Materials: Tubes (small or medium size), watch glasses, crucible* and cover*, crucible pliers, clay triangle. Safety Be careful when heating the crucible and lid. Hot crucible looks like cold, avoid direct contact with crucible, clay triangle and ring stand until you make sure they are cooled. When cleaning the crucible with concentrated nitric acid, use crucible pliers. In this section we will demonstrate dehydration and rehydration of chloride hexahydrate (II). When gently heated, the red burgundy \(\ce{CoCl2*6H2O}\) decomposes into violet \(\ce{CoCl2*2H2O}\) and then into blue anhydrous \(\ce{CoCl2}\). When this anhydrous compound dissolves in water, it will return to its original red burgundy color. In the dish, gently evaporate a small amount (0.3 – 0.5 g) of \(\ce{CoCl2*6H2O}\ crystals until purple and then blue. When this color change appears complete, add 3 to 5 ml of water and observe the color of the dissolved substance. Then heat the solution to dryness. In this section, you can observe changes in the physical properties of compounds, including humidity, color, structure, texture, and mass. Then it is necessary to decide whether the compound is hygroscopic, eruption or not using a change in the mass of the substance. On the analytical balance, weigh a pea-sized sample of each of the compounds below on separate clean and dry watch glasses. Record values as initial masses of containers and samples. Mark and place all the samples in the same place in the room so that they are not spilled. After an hour, pay attention to any changes in the physical appearance of each sample. Weigh the samples and record the masses as final masses. Calculate the weight change for each sample. A substance is classified as eruption if its mass is reduced by 0,005 g or more; and is classified as hygroscopic if its mass increases by 0.005 g or more. Compounds tested: \(\ce{Na2SO4*10H2O}\), \(\ce{FeCl3}\), \(\ce{KAl(SO4)2}\), \(\ce{CaCl2}\), \(\ce{CuSO4}\). In this section we will try to determine by studying a series of compounds that are real hydrates. Some compounds may have some hydrate properties without real hydrates. In order for a compound to be a true hydrate, it must exhibit all the properties of a true hydrate, including the evolution of water after heating, the solubility of its anhydrous residues in water and the reversibility of the colour of the residue back to the hydrate colour after dissolution in water. For each of the chemical compounds below, place the pea-sized (≈30 mg) amount of the compound (enough to cover the bottom of the tube) in a dry tube and note its color. Heat the tube and note the condensation that may appear in the mouth of the tube as evidence of dehydration, pay attention to the color of the Allow the residue to cool (place the tube in a beaker not on a plastic tube rack), then try to dissolve in about 3 ml of water (about 1/3 of a small tube), gently heating if necessary dissolve the residue (dissolve only substances that have shown condensation). Pay attention to the color of the dissolved residue. If the compound has all three of the above-mentioned properties, it is a real hydrate; if at least one of them is not present, the relationship is not. Test compounds: nickel chloride (II), cobalt chloride (II), sacrose, calcium carbonate, baru chloride, sodium tetraborate, potassium chloride. In this section, we will determine the number of water moths present on the ahydrous mole of the solid in a given hydrating. Using crucible pliers, clean the porcelain crucible and its lid with concentrated nitric acid (6 M). Pour the used nitric acid into the supplied waste container. Rinse the crucible and its lid with distilled water. Place the crucible with the lid slightly open on the clay triangle and heat firmly for at least 10 minutes. Allow the crucible to cool to room temperature (do not set the hot crucible at the top of the bench). Weigh the crucible with the lid to the nearest 0,001 g. Making sure that the crucible and its lid are clean pliers, add about 1 g (weighed to the nearest 0,001 g) of the unknown hydrate. Record the weight of the crucible, lid and sample. Heat the contents of the crucible with the lid slightly open so that the moistened water escapes, first gently (about 10 minutes) and then firmly (about 5 minutes). Center the crucible cover and allow to cool to room temperature. Weigh and record the mass of the chilled crucible with the lid and contents (anhydrous residues). Save the residue and perform calculations. If the results of your calculations suggest that some water is left in the residue, heat the sample for an additional 5 minutes, let it cool down and weigh again. After the experiment, dissolve all heated residues in water, put all solids and liquids in waste crock. A student trying to determine whether a white lump is a real hydrate heats the sample and finds that there is an evolution of water, that the resulting residue is soluble in water and that the solution is colorless. Is it a real hydrate? Explain. 1 The student receives a cobalt chloride hydrate (II). It weighs a clean and dry crucible with a lid and records a weight of 18,456 g. It then weighs the sample in a crucible and covers it and obtains a weight of 19,566 g. It heats the sample, allows it to cool to room temperature and reloads it to obtain a mass of 19,062 g. In this process, the color of the sample has changed from red burgundy to blue. Hydrate weight: Anhydrous mass \(\ce{CoCl2}\): Water mass driven away: Water moths: Anhydrous moles \(\ce{CoCl2}\): Moles per mole \(\ce{CoCl2}\): Hydrate formula: Why has the hydrate color changed? What color can you expect when this student dissolves the blue remnants in the water at the end of the experiment? Record your observations: Substance Initial mass of container and sample End weight of container and sample Change weight Observations regarding structure, texture, humidity, etc. Request 1 \(\ce{CaCl2}\) 2 \(\ce{Na2SO4*10H2O}\) 3 \(\ce{KAl(SO4)2*1 2H2O}\) 4 \(\ce{CuSO4}\) 5 \(\ce{FeCl3{FeCl3}\) Substance Initial color Water after heating (Y/N Colour (Y/N) Residue Soluble residue (Y/N) Colour of dissolved hydrate residue (Y/N) 1 Nickel (II) chloride 2 Cobalt chloride (II) 3 Sacrose 4 Calcium carbonate 5 Baru chloride 6 Sodium tetraborate 7 Potassium chloride Date Crucible weight and cover: Crucible weight, lid and solid hydrate: Crucrio mass, cover and anhydrous solid: Calculations Hydrate mass: Mass of the anhydrous solid: Water weight lost: Formula of anhydrous solid substance (from instructor): Molar mass of anhydrous mass: Moles \(\ce{H2O}\) present in hydrate: Moles of anhydrous solid present: Ratio of moles \(\ce{H2O}\) :Anhydrous lump = \(x\) : Hydrate formula [\(\text{Anhydrous solid}\ce{*}x\ce{H2O}\)]: Name of this compound: Unknown id: Did the compound(s) that appeared wet in Section B lost or gained water? Explain what might have happened. What will be the effect, on the mass of residue, overheating of the hydrate, so that the compound decomposes. Is this likely to lead to a higher or lower \(x\) value than the actual value? What will be the effect, on the mass of residue, not to heat the hydration enough to drive away all the water hydration in hydration. Is this likely to lead to a higher or lower \(x\) value than the actual value? Value?

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