Copper sulfate use and disposal

Posted by Anonymous on Fri, 2018-05-04 18:10

Copper (II) sulfate use and disposal

What is the equation for making solutions of copper (II) sulfate from solid and can you reuse copper sulfate crystals to make a saturated solution?

Can copper sulfate be used for crystal making or is there a safer and cheaper alternative?

Can students conduct a colour change reaction using copper sulfate solution, 6M HCl and ammonia or should this be conducted as a demonstration?

Is it safe to heat copper sulfate solid as an alternative to barium chloride to determine empirical formula or is there a safer alternative?

Voting:
☆☆☆☆☆
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Laboratory Technicians:
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Copper sulfate waste

Submitted by sat on 04 May 2018

Answers reviewed 27 February 2023

Science ASSIST recommends a site-specific risk assessment be conducted prior to the use of any chemicals taking into consideration:

- The use of safe operating procedures
- The equipment and how it will be used
- The chemical(s) used, their concentration and how they will be used
- By-products, chemical waste and any other waste produced
- The students' skill-level, behaviour, and ability to follow instructions

We have developed a one-page risk assessment template. https://assist.asta.edu.au/resource/2298/risk-assessment-template

Preparation of a copper sulfate solution of known concentration

When making a chemical solution, the concentration can be expressed as a percentage (eg. 5g/100ml is 5%), or as a molar concentration or molarity. Molarity is defined as the number of moles of solute dissolved in one litre of solution, and has the unit mol/L, denoted by upper case M. When preparing a solution where the concentration is expressed as a molarity, the formula to use is

m = c x V x M

where:

m = mass of solute (in grams)

c = concentration of solute (in moles per litre)

V = volume (in litres)

M = molecular weight (or molar mass) of solute (in grams)

The molecular weight can be found on the bottle label, or in a data book or safety data sheet (SDS), or by adding together the atomic weights of all the atoms which appear in the chemical formula of the substance. If you are using a hydrated salt, the water(s) of hydration must be included in the calculation of the molecular weight.

Example: Preparation of 1 litre of a 0.5 M copper sulfate solution

Starting with copper sulfate pentahydrate, ($CuSO_4.5H_2O$), we have

Molecular weight = M = 63.55 + 32.06 + (4x15.99) + 5x((2x1.008) + 15.99) = 249.68 g

Concentration = c = 0.5 M

Volume = V = 1 L

The quantity of the solid copper sulfate pentahydrate required is therefore

m = c x V x M

m = 0.5 x 1 x 249.68

m = 124.84 g

The copper sulfate should be added to about half to two-thirds of the final solution volume (about 500–650 mL) of distilled water in a large beaker, and stirred until dissolved. Gentle heating will speed up the dissolution of the salt. The solution can then be transferred to a measuring cylinder or volumetric flask and distilled water added to make up the volume to 1 litre.

Solutions of copper sulfate can degrade due to hydrolysis of the copper (II) ion with the formation of a precipitate of copper (II) hydroxide ($Cu(OH)_2$). and indicated by a cloudy solution. The copper sulfate solution can be stabilised by addition of a 0.1 M sulfuric acid solution in small aliquots (1–5 mL) until the solution becomes clear.

The heating of copper sulphate in solution is relatively safe. Hazards arise when heating the solid chemical. When copper sulphate is heated to decomposition (above ~250 °C), irritating copper oxide fumes and corrosive sulphur trioxide gases are produced. The temperatures involved in crystal making are less than 100 °C, which is nowhere near hot enough to cause decomposition.

To prevent heating to boiling point Science ASSIST recommends using hotplates or water baths, where possible, to prepare solutions. Bunsen burner temperatures are in excess of that required for crystal making. Rapid evaporation from copper sulfate solutions can result in hazardous fumes or spitting solids. The work area should be well-ventilated.

Method to prepare a saturated solution of copper sulfate using copper sulfate crystals

Copper sulfate crystals made by students can be reused to make new copper sulfate solutions.

- Ensure the working area is well ventilated.
- The solubility of copper sulfate pentahydrate is 317g/L water at 200C. It is recommended that you weigh the copper sulfate crystals to determine the approximate volume of water required to make a saturated solution.
- Small crystals of copper sulfate (less than 5mm in size) can be placed directly into deionised water and stirred to dissolve.
- Large crystals can be crushed gently in a mortar and pestle to smaller crystals to enable better dissolving. Alternatively larger crystals can be allowed to dissolve in sufficient water over several hours.
- Where an inhalation risk exists, work in a fume cupboard or consider wearing a Class P1 (particulate) respirator. See Australian/New Zealand Standards AS/NZS 1715 and 1716 for more information1 , refer to the SDS from your copper sulfate supplier.
- Heating the copper sulfate solution to a temperature of around 50-600C will speed up dissolution. The use of a hot plate or water bath is appropriate as long as the temperature is monitored.
- Never inhale any fumes.
- If required, filter the solution to remove any undissolved solid and any impurity that may be present from using technical grade chemicals or dust or other contaminants from the environment during the crystal making process.
- Solutions made from recycled copper sulfate crystals would not be suitable for senior analytical experiments.

Growing copper sulfate crystals

Copper sulfate is commonly used as a fungicide, herbicide, and pesticide. It will produce striking, blue, diamond-shaped crystals in the laboratory.[2, 3] Using chemicals to grow crystals is a common activity in school science laboratories. Crystals are solids that form when molecules join together in a regular repeating pattern. To grow crystals, a saturated solution is required and the method is outlined in the references below.[4]

Note: Some technical grade copper sulfate chemicals may contain impurities, which will affect the formation of crystals and the solution will need to be filtered when recycling for reuse. For more information, refer to the Flinn Scientific Laboratory Solution Preparation reference.[5]
https://www.flinnsci.com/laboratory-solution-preparation/dcat016/

Alternative chemicals for growing crystals

Aluminium potassium sulfate (potash or common alum) is found in industrial, medicinal, cosmetic and food products and is widely used as a preservative in food processing. It can produce large clear crystals in the laboratory.[6]

The solubility of aluminium potassium sulfate increases significantly with temperature. At 25 °C, the solubility is about 14g/100mL, whereas at 100 °C, this increases to about 160g/100mL. Therefore, 100 mL of hot, saturated solution may cool to give an excess of over 140g of alum, which would be unlikely to precipitate as large, well-formed crystals.

Both copper sulfate and aluminium potassium sulfate are considered safe to use with students from Years 7–12,[7] if proper controls are put into place to deal with the hazards. Students should wear safety glasses and avoid skin contact. Hands should be washed immediately after completing the activity and students should not be allowed to take the crystals home. It is recommended that any crystals formed are recycled for future use. Sugar and salt could also be used as a safe alternative.

Using copper sulfate solution in a colour change experiment

Reacting copper sulfate with hydrochloric acid produces a colour change, which is reversible with the addition of ammonia.[8] Science ASSIST recommends that 6M HCl be used by the teacher as a **demonstration** only and not by the students. Science ASSIST considers that up to 2M HCl is sufficient for most general chemistry activities and recommends using the lowest concentration that delivers the desired outcome.

Using copper sulfate to determine empirical formula

Heating hydrated copper sulphate is a useful activity to demonstrate both chemical and colour changes. However, it is difficult to remove all of the water of crystallisation by heating without also decomposing some of the copper sulfate. Therefore, it is not recommended to heat copper sulfate for determining an empirical formula.

Dehydration of copper sulfate pentahydrate

When copper sulfate pentahydrate is heated, the loss of water of crystallisation proceeds in a stepwise manner.[9] From 30 to 110°C, it loses two molecules of water to give the trihydrate, followed by the loss of two more water molecules to give the pale blue monohydrate. Further heating above 150°C gives white anhydrous copper sulfate.

$CuSO_4.5H_2O_{(s)} \rightarrow CuSO_4.3H_2O_{(s)} + 2H_2O_{(g)}$

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$CuSO_4.H_2O_{(s)} \rightarrow CuSO_{4(s)} + H_2O_{(g)}$

Decomposition of copper sulfate occurs at about 650°C. The products of decomposition are black copper (II) oxide and sulphur trioxide:

$CuSO_{4(s)} \rightarrow CuO_{(s)} + SO_{3(g)}$

Sulphur trioxide is a highly corrosive substance, which is irritating to the respiratory system and combines with water to form sulphuric acid. The dehydration of copper sulfate pentahydrate is more suited to a qualitative demonstration of a reversible reaction, where complete dehydration of the salt is not required.

Dehydration of barium chloride dihydrate

Hydrated barium chloride will not decompose under laboratory heating conditions and does not change colour. Barium chloride dihydrate loses its water of crystallisation when heated to 120°C. Its melting point is greater than 900°C and it therefore can withstand heating over a Bunsen flame. It is, however, more toxic than copper sulfate and its dehydration does not result in a colour change.[10]

Alternative reagents suited to empirical determination experiments

Oxidation of magnesium ribbon to magnesium oxide

Magnesium oxide can be generated from magnesium ribbon by heating the magnesium in a crucible over a Bunsen flame.[11] The crucible is placed in a pipe clay triangle (with the lid slightly off-centre in order to expose the metal to air) and is heated until the contents turn grey-white, indicating the transformation of the magnesium to the oxide. The empirical formula of magnesium oxide is then calculated from the difference in the initial mass of the magnesium ribbon and that of the magnesium oxide produced.

An environmentally friendly alternative[12]

Magnesium sulfate or Epsom salts offer a safe alternative for determining an empirical formula. Although there is no colour change, there are physical differences in the dried chemical.

Safe handling, recycling and disposal of copper sulfate

- Copper sulfate is a skin and serious eye irritant and is toxic by ingestion and inhalation. Avoid contact with eyes and skin and breathing dust.
- Wear appropriate PPE such as safety glasses and nitrile or latex gloves.
- Do not dispose solutions down the sink, as copper sulfate is toxic to aquatic life.
- Waste solution treatment options:
- evaporate water from the solution and store residue for collection by a waste contractor
- displace copper ions by adding in steel wool. Solid copper can be disposed of in the regular waste bin, the resulting clear solution can be washed down the sink.
- Recrystallise the salt and recycle
- Precipitate the copper ions by adding the solution to a slight excess of sodium carbonate solution. Collect precipitate by filtration for collection by a waste contractor, the filtrate can be put down the sink.

Copper sulfate poisoning

According to a recent study,[13] copper sulfate was responsible for 55 cases of poisoning in science classes, 45% of which were hospitalised.

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