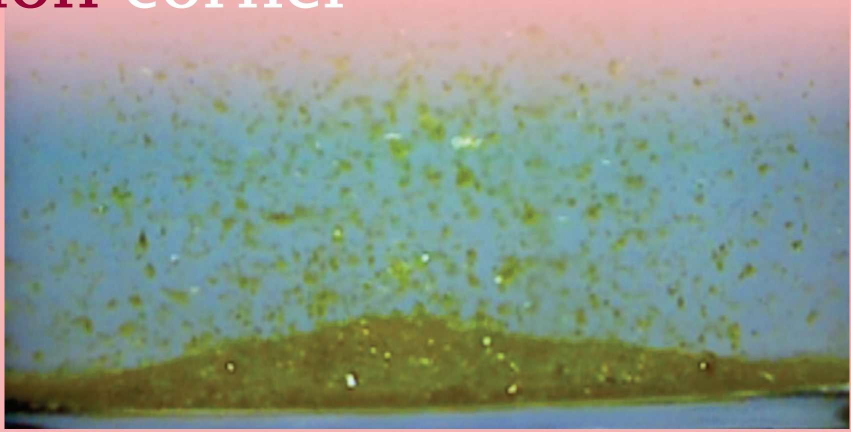
Demonstration corner

GOLDEN RAIN

This is a beautiful demonstration that is a lovely introduction to discussions about precipitation and solubility. If you want to do this on a larger, or smaller scale, just vary the quantities accordingly.



You will need

- 0.3% solution of lead nitrate (0.3 g in 100 cm^3 of water)
- 0.3% solution of potassium iodide (0.3 g in 100 cm^3 of water)
- A few drops of 1 M hydrochloric acid
- 250 cm^3 conical flask
- Dropping pipette
- Kettle and large beaker for a water bath or a hotplate.

Health & Safety

This is not a dangerous demonstration but it does involve the use of lead compounds, albeit at low concentrations: Wash hands thoroughly after the demonstration. Wipe up any spills and wipe over surfaces.

The demonstration

1) Prepare the solutions by dissolving the solids each in 100 cm^3 of distilled water.

   If you look at the stoichiometry, you will see that the potassium iodide is in excess: this is to maximise the chances of precipitating lead ions out of solution and reducing the possibility of washing away dissolved lead during disposal. When you have made the solutions, add a few drops of 1 M hydrochloric acid to each: this is to prevent the formation of lead carbonate which can be formed in impure water (including distilled water that has stood long enough for sufficient CO_2 to dissolve in it). Lead carbonate has a very low solubility and the haziness of its precipitate can ruin the effect.

2) Pour one of the solutions into the other. You will immediately get a brilliant yellow precipitate of lead iodide

   $$Pb(NO_3)_{2(aq)} + 2KI_{(aq)} \longrightarrow 2KNO_{3(aq)} + PbI_{2(s)}$$

   The tiny crystals of lead iodide that form swirl around in the flask and the concentration gradients as the liquids mix combine to generate a 'pearlescent' effect similar to that you might see in some shampoos or similar liquids.

3) Now heat the solution slowly, either in a water bath or on a hotplate.

   Whilst the lead iodide may be insoluble in water at room temperature, its solubility increases slightly with temperature. Put simply, when ionic compounds dissolve in water, they dissociate into their component ions. This dissociation can either give out energy (exothermic) or take in energy from the surroundings (endothermic), depending on the substance. In the case of lead iodide, it dissociates into Pb^{2+} and I^- ions. This is endothermic, and so increasing the temperature of the solution will promote the dissociation of lead(II) iodide. Consequently, the solubility of lead iodide rises from 0.0756 g per 100 cm^3 of water, to a still not very impressive 0.19 g per 100 cm^3 of water.

   The concentrations of the solutions are such that by the time the mixture reaches 60°C, all the lead iodide will have dissolved and the solution will be colourless again.

4) Leave the solution to cool, somewhere you, and your class can see it.

As the solution cools, the dropping temperature forces very pure crystals of lead iodide to precipitate back out of solution. These hexagonal crystals are larger than the very fine particles formed initially but they still take some time to meander gently to the bottom of the flask, giving the reaction mixture a shimmering, glittering effect commonly referred to as a 'golden rain'. The effect can last for up to an hour as the crystals fall out of the solution.

The effect is much better if you shine quite a bright light onto the flask. As the crystals tumble out of solution they reflect the light in tiny sparkles.

Disposal

Even though the iodide is in excess, there will still be some free lead ions in solution. Add a few cm^3 of 1 M sodium carbonate solution to precipitate out any remaining ions as lead carbonate (for this volume 2 cm^3 should be plenty). Filter the precipitates out and keep for disposal. The filtrate can be washed to waste with plenty of cold running water.