Determination of a Solubility Product Constant

EXPERIMENT 19C

When an ionic solid dissolves in water, it dissociates to give the positive and negative ions that make up the solid. These ions are hydrated, and are found in solution in the same relative proportion as in the solid. As more solid dissolves, the concentration of the ions increases. This build-up allows the reverse reaction, in which the ions crystallize out, to have a greater possibility of occurring. Eventually, a situation is reached in which the rate of dissolving is equal to the rate of crystallization. At this point, no more solid can dissolve, and the solution is said to be *saturated*. A state of equilibrium has been reached which can be recognized by a constant color for the solution (if it is colored), or by a constant mass left undissolved.

The solubility product constant, K_{sp}, for an ionic solid is given by the product of the concentration of the ions, each raised to the power of the coefficients in the dissolving reaction. For instance, the K_{sp} for silver chloride, AgCl, is given by

$$K_{sp} = [Ag^+][Cl^-]$$

For a substance such as lead iodide, PbI_2, the K_{sp} is given by

$$K_{sp} = [Pb^{2+}][I^-]^2$$

The K_{sp} expression gives the relationship between the ions in the saturated solution, and therefore their maximum possible concentration without causing precipitation. If solutions of suitable concentration of substances are available, it is possible to mix them to form a precipitate, then carry out appropriate dilutions until a point is reached at which no precipitate occurs. This process allows an approximate value to be determined for the K_{sp}.

In this experiment, you will mix solutions of lead nitrate and potassium iodide at a number of different dilutions, and watch for the first situation in which no precipitate occurs. You will then be able to state the K_{sp} at room temperature as a range of values. The test tubes in which a precipitate did occur will then be heated until the precipitate dissolves, in order that you may determine the K_{sp} at different temperatures.

OBJECTIVES

1. to prepare a number of solutions of each of Pb^{2+} and I^-, of differing concentrations
2. to mix combinations of the above solutions and note whether a precipitate occurs
3. to obtain an approximate value of the K_{sp} for PbI_2 at room temperature
4. to obtain the approximate K_{sp} for PbI_2 at temperatures higher than room temperature

MATERIALS

Apparatus

12 test tubes (18 mm × 150 mm)
2 test-tube racks
2 graduated cylinders (10 mL)
medicine dropper
beaker (400 mL)
2 beakers (100 mL)
laboratory burner
ring stand and ring
wire gauze
thermometer
water soluble marker
lab apron
safety goggles

Reagents

0.010*M* $Pb(NO_3)_2$
0.020*M* KI

PROCEDURE

CAUTION: Lead nitrate is very toxic. Do not get any in your mouth and do not swallow any. Avoid getting any on your skin, since it can be absorbed. Wash away any spills or splashes with plenty of water. Call your teacher.

1. Put on your lab apron and safety goggles.
2. Obtain in separate 100 mL beakers about 40 mL of each of 0.010*M* $Pb(NO_3)_2$ and 0.020*M* KI, and label the beakers.
3. Obtain twelve 18 mm × 150 mm test tubes and arrange them in two racks, each with 6 test tubes. Label each set *A* to *F* with your water soluble marker. Label them near the top, since they will be immersed in water later.
4. Into the first set of test tubes place 10.0 mL, 8.0 mL, 6.0 mL, 4.0 mL, 3.0 mL, and 2.0 mL of 0.010*M* $Pb(NO_3)_2$, respectively. Use your 10 mL graduated cylinder, and get the precise amount by adding or subtracting drops with a dropper.
5. Add an amount of water to each tube to make the volume in each up to 10.0 mL (that is, 0.0 mL, 2.0 mL, 4.0 mL, 6.0 mL, 7.0 mL, and 8.0 mL, respectively).
6. Repeat Steps 4 and 5 with the test tubes in the second rack, using 0.020*M* KI instead of 0.010*M* $Pb(NO_3)_2$.
7. Mix the contents of test tube *A* from the lead nitrate set with the contents of test tube *A* from the potassium iodide set, and replace the test tube in the rack.
8. Repeat Step 7 for each of the other 5 combinations.
9. Record in which test tubes a precipitate occurs in your copy of Table 1 in your notebook.
10. Add about 250 mL of water to a 400 mL beaker, and place the beaker on a wire gauze on a ring clamp attached to a stand.
11. Place each of the test tubes which contain a precipitate in the beaker, and begin heating slowly with a laboratory burner. (See Figure 19C-1.)
12. When the precipitate in each test tube dissolves, note the temperature of the water bath and record it in Table 1.
13. Before leaving the laboratory, wash your hands thoroughly with soap and water; use a fingernail brush to clean under your fingernails.

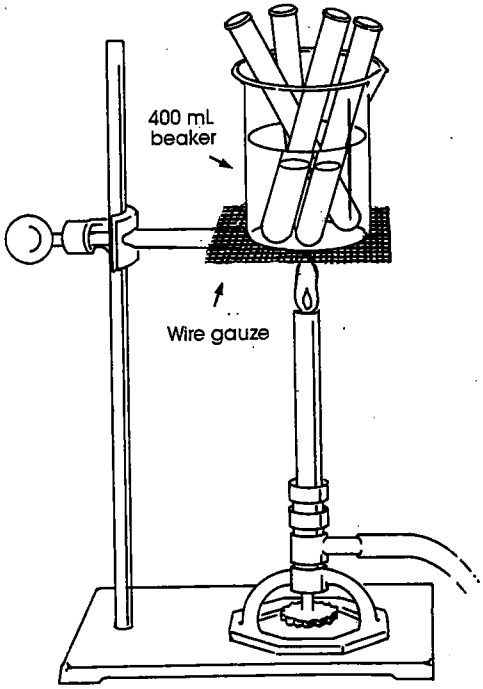

Figure 19C-1