Experiment 5

Lowering of the Freezing Point

Purpose :

To determine the freezing point of a solvent and of dilute solutions of various known concentrations, and to calculate the molecular weight of the solute from freezing-point lowering.

Principle :

When a solute is dissolved in a solvent, a lowering of freezing point occurs which is proportional to the effective molecular concentration of the solute. If the freezing points are carefully determined for a pure solvent and for a dilute solution containing a weighed quantity of solute dissolved in a known amount of solvent, the molecular weight of the solute may be calculated from the following equation:

$$M = \frac{1000 K_f g}{\Delta T_f G} \tag{1}$$

where M is the molecular weight of the solute,

K_f is the molal freezing-point depression constant,

g is the weight of the solute,

G is the weight of solvent used,

ΔT_f is the lowering of the freezing point.

Accurate measurement of the freezing-point lowering is made possible by the use of the differential thermometer devised by **Beckmann**. The scale of this thermometer is divided into tenths and hundredths. The total range of the scale is five or six degrees. The thermometer is adjusted to read on scale in the desired temperature range by varying the amount of mercury in the large lower bulb. The reservoir at the top of the Beckmann thermometer is used to receive the excess mercury from, or to provide additional mercury to, the lower reservoir in "setting" the Beckmann thermometer to read on scale in the desired temperature range.

When the solute is an electrolyte, the effective molecular concentration will be greater than that calculated from the moles of electrolyte added to the quantity of solvent used, because of the dissociation of the solute into ions. The "apparent" molecular weight as determined from freezing point measurements on electrolyte solutions may be used to calculate the degree of ionization α, if the solute is a weak electrolyte. If i represents the true molecular weight divided by the "apparent" molecular weight, then

$$\alpha = \frac{i-1}{n-1} \tag{2}$$

where n is the number of ions formed by the dissociation of one molecule of the electrolyte used.

Apparatus and Chemicals:

Beckmann thermometer with rubber stopper; inner tube; air jacket; safety bulb; *25-ml* pipette; spatula; *600-ml* beaker; $100°C$ thermometer; glass bar; ice vessel.

NaCl, ice and NaCl for producing a freezing mixture.

Procedures :

(1) Prepare a freezing bath which is about $5°$ below the freezing point of the solvent been tested.

(2) Pipette accurately *25 ml* distilled water into inner tube. Setup the Beckmann thermometer. Be sure it is clean and dry.

(3) The liquid is cooled until it is approximately $1°$ above its freezing point by immersing the inner tube directly into the freezing bath. Then the tube is thoroughly dried and place in the air jacket.

(4) Assemble the apparatus as **Fig. 5**.

(5) Stir gently and continuously. Tap the stem of the thermometer at regular intervals and record the temperature reading every *10 sec*. When the equilibrium temperature has been located, record 20 more temperature readings. The value obtained when readings have become constant is taken as the freezing point of pure solvent.

(6) Take out the tube and melt the solvent by hand. Add accurately four successive additional weighed portions of *0.15 g* of NaCl and repeat **steps (4)-(5)** to determine the freezing point for each concentration.

Calculations :

(1) Draw cooling curves on coordinate paper for each run and determine the freezing points.

(2) Calculate the molecular weight (or the apparent molecular weight) of the solute at each dilution used.

(3) Plot the molecular weights obtained against added NaCl's weights. When this curve is extrapolated to zero concentration, get the experimental molecular weight.

(4) If the solute is an electrolyte, calculate the degree of dissociation for each dilution and plot against added NaCl's weights.

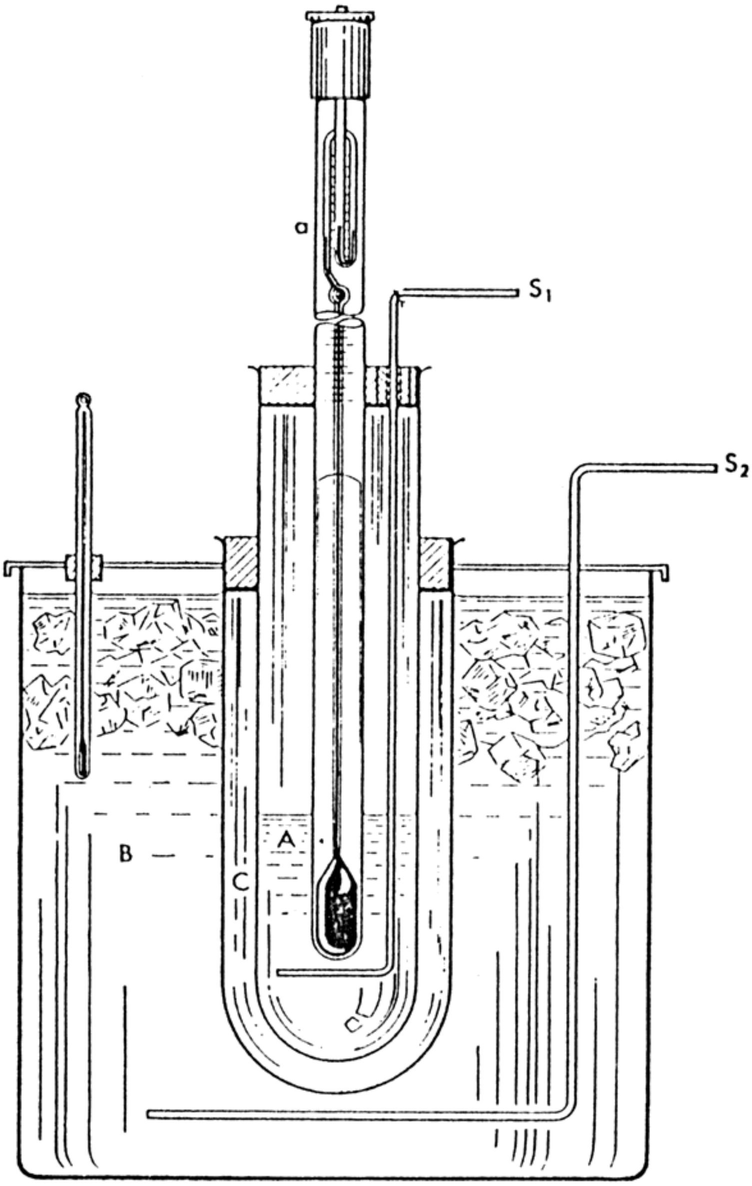

Fig. 5

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