

TEACHING WATER ANALYSIS

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In the commercial laboratory, water analysis is one of the more important of the various analytical determinations a chemist is called on to do. It follows, therefore, that in classes of quantitative analysis, instruction in water analysis should be included. Otherwise the graduate seeking a position as chemist is likely to find himself unqualified.

The main difficulty in teaching students this determination lies in the fact that analyzed samples are not available to the average instructor. Few college chemistry departments can afford to have several water samples analyzed each year. The alternative course is for the instructor to analyze his samples himself. This is a great burden on the instructor and he is apt instead to omit water analysis instruction altogether. The writer believes he has solved this problem and offers this suggestion:

Make up molar solutions of the following salts:

MgSO₄, CaCl₂, NaCl, Mg (NO₃)₂, NaNO₃, and NaHCO₃.

(These salts — or their hydrates — are obtainable in definite and pure form. Solutions of them at one molar concentrations are easily made by weighing out the correct quantities on an analytical balance. These will be near enough to 1.0 M for the purpose at hand provided the finest grade salts are used.)

The instructor is now ready to give out water samples to his students and should ask them each to bring clean, dry, gallon bottles for the purpose. It is suggested that all samples be given out on a single afternoon. The reason will be apparent in what is to follow).

Fill a two-liter volumetric flask with tap water. (This, being the stock solution for all samples, must be analyzed. The instructor can do this himself, or give it to a couple of his better students as their unknowns. In the latter case, it is well for the instructor to check their analyses by mak-

ing a total solids determination himself.) The hardness of this two-liter stock sample may now be increased by adding exactly 1 ml portions of one or more of the salt solutions. The increase in hardness is easily calculated, as in the following example:

Suppose, in preparing a sample, 1 ml of 1.0 M CaCl₂ and 1 ml of 1.0 M NaNO₃ were added to the 2 liters of stock solution of tap water. The hardness of the tap water would be increased by

$$\frac{1}{1000} \times \frac{1}{2} \times \frac{40.08}{1000} \times \frac{1000}{1000}, \quad \text{or} \quad 20.04 \text{ p. p. m. of Ca}^{++}$$

(a) (b) (c) (d)

- (a) Gram ions of Ca⁺⁺ added to 2 liters
 (b) Gram ions of Ca⁺⁺ per liter
 (c) Gram of Ca⁺⁺ per liter or parts per thousand
 (d) Parts per million of Ca⁺⁺

$$\frac{2}{1000} \times \frac{1}{2} \times \frac{35.457}{1000} \times \frac{1000}{1000}, \quad \text{or} \quad 35.457 \text{ p. p. m. or Cl}^-$$

(a) (b) (c) (d)

- (a) Gram ions of Cl₁ added to 2 liters
 (b) Gram ions of Cl₁ per liter
 (c) Grams of Cl₁ per liter or parts per thousand
 (d) Parts per million of Cl₁

and, by similar calculations, 11.50 p.p.m. of Na⁺
 and, 31.004 p.p.m. of NO₃

By adding various combinations of the salt solutions, the compositions of the samples can be varied over a wide range and yet their correct compositions remain known.

The following table may be used to simplify the arithmetic if the samples are made up as suggested above:

Salt Sol'n Used	Increase in Hardness of Stock Solution (Expressed as parts per million)
CaCl ₂	20.04 p.p.p.m. of Ca ⁺⁺ + 35.457 p.p.p.m. of Cl ⁻
MgSO ₄	12.16 p.p.p.m. of Mg ⁺⁺ + 48.03 p.p.p.m. of SO ₄ ⁻²
Mg (NO ₃) ₂	12.16 p.p.p.m. of Mg ⁺⁺ + 62.008 p.p.p.m. of NO ₃ ⁻

NaNO_3	11.50 p.p.m. of Na^+ + 31.004 p.p.m. of NO_3^-
NaCl	11.50 p.p.m. of Na^+ + 17.73 p.p.m. of Cl^-
NaHCO_3	11.50 p.p.m. of Na^+ + 30.51 p.p.m. of HCO_3^- (or 15.00 p.p.m. of CO_3^{2-})

It is the custom of many commercial chemists to report hardness in grains per gallon instead of parts per million. The latter is easily converted to the former by means of the factor 0.0584. Thus 30.0 p.p.m. is $30.0 \times .0584$, or 1.752 g.p.g.