Tikrit University
College of Veterinary Medicine

Lecture Title:
Preparation of Standard Solutions

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Preparation of Standard Solutions

A worker or operator in charge of food analysis must understand the methodology of preparing standard chemical solutions. The determination of dietary elements depends on performing chemical titrations between standard solutions to calculate the percentage of nutritional compounds in the sample, using specific equations and laws.

Standard Solution

A standard solution is prepared by dissolving the equivalent weight of a chemical substance in one liter of distilled water.

- **Equivalent weight** of a substance =

Molecular weight÷Valence

- **Molecular weight** = summation of the atomic weights of all atoms in the compound.
- **Valence** = number of hydrogen atoms that can combine with, or be replaced by, the substance.

Thus:

Equivalent weight = $\frac{Number\ of\ atoms \times Atomic\ weight\ (at.wt.)}{Number\ of\ hydrogen\ atoms\ (pH)}$

Table (1): Atomic weights (at.wt.) and atomic numbers (at.No.) of related elements

Element	Symbol	At. wt.	At. No.
Aluminum	Al	26	13
Arsenic	As	74	33
Barium	Ba	137	56
Calcium	Ca	40	20
Carbon	C	12	6
Chlorine	Cl	35	17
Chromium	Cr	52	24
Cobalt	Co	58	27
Copper	Cu	63	29
Hydrogen	H	1	1
Iodine	I	126	53
Iron	Fe	55	26
Lead	Pb	207	82
Magnesium	Mg	24	12
Manganese	Mn	54	25
Mercury	Hg	200	80
Molybdenum	Mo	95	42
Nickel	Ni	58	28

Nitrogen	N	14	7
Oxygen	O	16	8
Phosphorus	P	30	15
Potassium	K	39	19
Silver	Ag	107	47
Sodium	Na	23	11
Sulphur	S	32	16
Zinc	Zn	65	30

Example 1: Preparation of Sodium Hydroxide Standard Solution (NaOH)

- Atomic weights: H = 1, O = 16, Na = 23
- Molecular weight = 23 + 16 + 1 = **40 g/mol**
- Valence = 1
- Equivalent weight = 40 ÷ 1 = **40 g**

Therefore, dissolve **40 g of NaOH granules** in one liter of distilled water.

Example 2: Preparation of Sulphuric Acid Standard Solution (H_2SO_4)

- Atomic weights: H = 1, O = 16, S = 32
- Molecular weight = (2×1) + (32) + (4×16) = 98 g/mol
- Valence = 2
- Equivalent weight = 98 ÷ 2 = **49 g**

Thus, dissolve **49 g of H_2SO_4** in one liter of distilled water.

If concentrated sulphuric acid is used (density ≈ 1.8 g/mL):

Volume= $\frac{Equivalent\ weightDensity}{Density}$

$= \frac{49}{1.8} = 27.2$ mL

So, measure **27.2 mL of concentrated H_2SO_4**, then dilute to one liter with distilled water.

Important note: Always pour acid into water (not the reverse) to prevent violent reactions and splashing.

Characteristics of Standard Solutions

Exact standard solutions react with each other in equal volumes. However, they are not commonly used directly in titration because:

1. They are highly concentrated.
2. Neutralization occurs too rapidly.
3. Small errors cause large deviations in results.

Therefore, **decinormal (1/10 N) solutions** are commonly used in practice because:

1. They economize chemical use.
2. They improve titration accuracy.
3. Small errors have minimal impact on results.

Normality (N)

Normality is the ratio between the actual dissolved weight and the theoretical equivalent weight:

$$N = \frac{Actually\ dissolved\ weight}{Equivalent\ weight}$$

In titration:

Volume of Acid×Normality of Acid=Volume of Base×Normality of Base

Indicators

Indicators are weak organic acids or bases that change color within specific pH ranges.

- **Methyl orange**: red in acidic medium, yellow in alkaline.
- **Methyl red**: also changes from red (acidic) to yellow (alkaline).

The **endpoint (balance point)** occurs when neutralization coincides with a distinct color change.

Worked Example: Normality of Sulphuric Acid Using Sodium Bicarbonate ($NaHCO_3$)

1. **Equivalent weight of $NaHCO_3$:**

$(3\times16)+(12)+(1)+(23)=84$ g

2. **Tenth-normal solution:**

   $$\frac{84}{10} = 8.4\ g/L$$

3. Prepare 0.84 g in 100 mL of distilled water.
4. Actual dissolved = 0.82 g →

$N=\frac{0.82}{0.84} = 0.96$

5. 20 mL of this base solution was titrated with H_2SO_4, giving 21.4 mL acid:

$21.4 \times Nacid = 20 \times 0.96$

$Nacid = \frac{19.2}{21.4} \approx 0.90$

Example: Normality of NaOH (Indirect Method)

- Step 1: Determine H_2SO_4 normality using $NaHCO_3$.
- Step 2: Use this acid to titrate NaOH.

Given:

- H_2SO_4 (titrated with $NaHCO_3$) = 11 mL
- NaOH (titrated with same acid) = 12.2 mL
- 10 mL of each base was taken.

Solution:

$11 \times N_{acid} = 10 \times 1 \Rightarrow N_{acid} = 0.91$

$12.2 \times 0.91 = 10 \times N_{NaOH}$

$N_{NaOH} = \frac{11.1}{10} = 1.11$