Determination of Calcium in Milk
Calcium is an important component of a healthy diet and a **mineral** necessary for life.

- It is a mineral that people need to build and maintain **strong bones and teeth**.

- It is also very important for other physical functions, such as **muscle control and blood circulation**.
If we do not have enough calcium in our diets to keep our bodies functioning, calcium is removed from where it is stored in our bones. Over time, this causes our bones to grow weaker and may lead to osteoporosis (a disorder in which bones become very fragile).
Milk and calcium:

- Milk is a heterogeneous mixture of proteins, sugar, fat, vitamins and minerals.
- Milk and milk products are some of the natural sources of calcium.
- Cow’s milk has good bioavailability of calcium (about 30 to 35%).

- Milk is an excellent source of dietary calcium for those whose bodies tolerate it because it has a high concentration of calcium and the calcium in milk is excellently absorbed.
Practical Part
Objective:

- Determination of Calcium in milk sample.
In this experiment, the determination of calcium in milk is based on a complexometric titration of calcium with an aqueous solution of the disodium salt of EDTA at high pH value (12).

**Complexometric titration** is a type of titration based on complex formation between the analyte and titrant.

Such compounds are capable of forming chelate complexes with many cations in which the cation is bound in a ring structure.

The ring results from the formation of a salt-like bond between the cation and the carboxyl groups together with a coordinate bond through the lone pair of electrons of the nitrogen atom.
Principle cont’:

- The common form of the agent is disodium salt Na$_2$H$_2$EDTA.
- It is colorless and can be weighed and dissolve in water to form a stable solution.
- At high pH (> 10) the remaining protons leave EDTA forming EDTA$^{4-}$ anion:
Indicator-Solochrome dark blue:

• The Solochrome dark blue indicator is a suitable indicator in this case.
• The dye itself has a blue color.
• This blue dye also forms a complex with the calcium ions changing colour from blue to pink/red in the process, but the dye–metal ion complex is less stable than the EDTA–metal ion complex.
• As a result, when the calcium ion–dye complex is titrated with EDTA the Ca2+ ions react to form a stronger complex with the EDTA changing the dye color to blue.

• Ca-Indicator + EDTA$^{4-}$ → Ca-EDTA$^{2-}$ + Indicator
Excess Ca\(^{2+}\) ions present to complex with indicator

Ca\(^{2+}\) ions almost all complexed by EDTA

All Ca\(^{2+}\) ions complexed by EDTA, indicator completely uncomplexed
Method:

- Combine 10mL of sample, 40mL distilled water, and 4mL of 8M sodium hydroxide solution into an Erlenmeyer flask and allow solution to stand for about 5 minutes with occasional swirling.
- A small of magnesium hydroxide may precipitate during this time. Do not add the indicator until you have given this precipitate a chance to form.
- Then add 6 drops of the Solochrome dark blue solution.
- After that start to titrate with EDTA solution.
- Repeat titration for three trials.
### Results:

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<tr>
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<th>EDTA volume (ml)</th>
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<tr>
<td>1</td>
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<tr>
<td>2</td>
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<tr>
<td>3</td>
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<td><strong>Average</strong></td>
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1. Calculate the moles of EDTA required to complex the Ca$^{2+}$ ions in the sample:

\[ \text{Number of moles (for EDTA)} = \text{Molarity of EDTA} \times \text{volume of EDTA in L} \]

**Note:** Ratio Ca$^{2+}$:EDTA = 1 : 1 (i.e moles of EDTA = moles of Ca$^{2+}$)

2. Calculate weight of Ca$^{2+}$:

\[ \text{Weight of Ca}^{2+} = \text{Number of moles} \times \text{molecular weight (40.78)} \]

\[ \text{% of Ca}^{2+} = \left( \frac{\text{weight of Ca}^{2+}}{\text{weight of sample}} \right) \times 100 \]

OR

\[ \text{Amount of calcium} = \left( \frac{\text{Molarity of EDTA} \times \text{vol. of EDTA (in liter)} \times 40.78}{\text{weight of sample}} \right) \times 100 \]