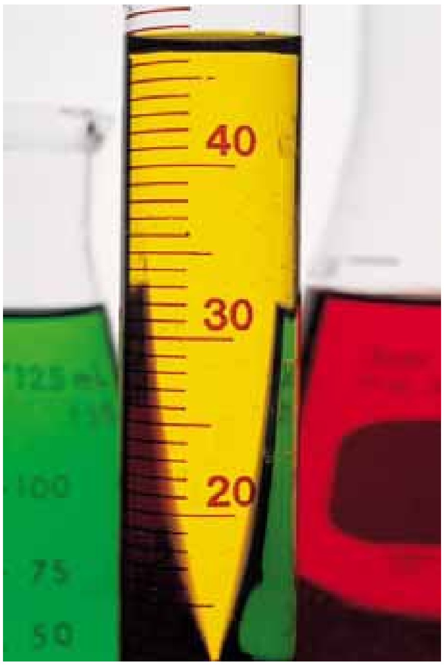
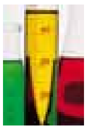
# Lecture 1: Precision of Chemical Measurements

June 26, 2004



# Uncertainty on a measurement

- **In every measurement there is a possibility of experimental error. The margin of error in a given measurement is called the** uncertainty **on that measurement. The uncertainty depends on the measuring tool or instrument.**
  - **Analytical balances:**

| Instrument | Uncertainty |
|---|---|
| **Mettler model PB 302** | ± **0.01 g** ← least precise |
| **Mettler model PE 160** | ± **0.001 g** |
| **Mettler model AB 104/204** | ± **0.0001 g** ← most precise |

# Analytical balance

Mass of beaker + water

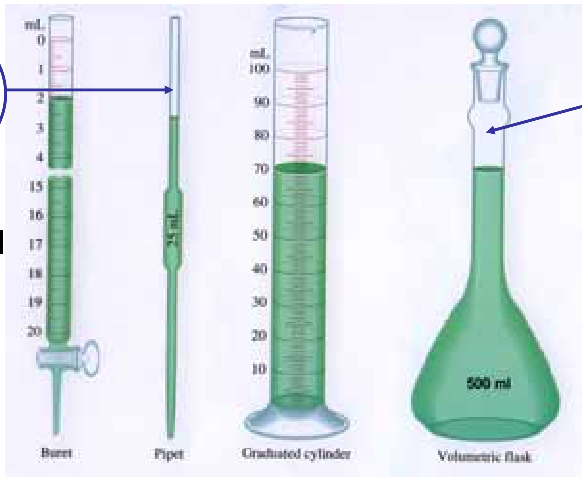
= 60.3997 g

# Uncertainty

– **Volumetric glassware:**

TD
±0.06 mL
25°C

**25.00 ml**




TC
±0.25 mL
25°C

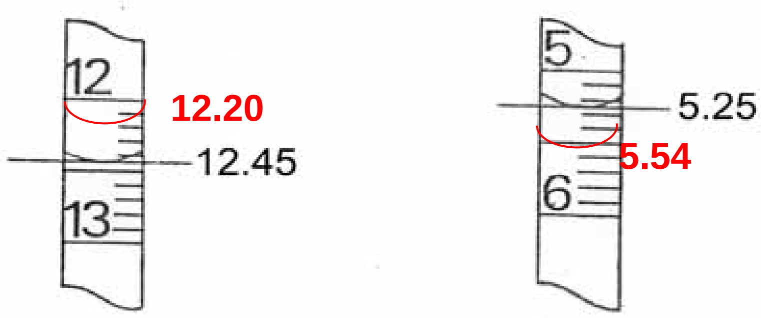
**500.0 ml**

# Reporting volume measurements

| Tool | Uncertainty | Correct volume writing |
|---|---|---|
| **10 mL graduated cylinder** | **± 0.1 mL** | **8.3 mL (for example)** |
| **50 mL buret** | **± 0.02 mL** | **24.85 mL (for example)** |
| **25 mL volumetric pipet** | **± 0.06 mL** | **25.00 mL** |
| **500 mL volumetric flask** | **± 0.25 mL** | **500.0 mL** |

# Buret reading

**Uncertainty = ± 0.02 or 0.03 mL**



# Significant figures

- Definition: **In a given number, each digit that appears to the right of the first non-zero digit, including that latter digit, is a significant digit (or significant figure:** Sig. Fig.**).**
- Examples:

| | |
|---|---|
| **0.0032** | ***2 Sig. Fig.*** |
| **1.2100** | ***5 Sig. Fig.*** |
| **5.00** | ***3 Sig. Fig.*** |
| **5** | ***1 Sig. Fig.*** |

- Note: **Zeroes to the left of the first non-zero digit and exponents are** NOT **significant.**

| | |
|---|---|
| $5.0 \times 10^{2}$ | ***2 Sig. Fig.*** |
| $500 \times 10^{6}$ | ***3 Sig. Fig.*** |
| $0.01 \times 10^{-4}$ | ***1 Sig. Fig.*** |

- Physical Significance: **A given number (measurement) contains significant figures that are all certain,** except one **(the last one) that is uncertain.**

# Addition and subtraction

- **The result must be expressed to the same degree of precision as that of the** *least precise* **component in the operation.**

```
  245.1      ←  least precise
+   6.713
   10.02
 --------
  261.833        Result: 261.8
       ↑
 we round off here
```

- Example from the Chemistry lab:

| | |
|---|---|
| **Upper reading of buret (mL)** | **13.35** |
| **Lower reading of buret (mL)** | **25.58** |
| **Volume** | **12.23** |

# Multiplication and division

- **Keep as many** *significant figures* **in the result as there are in the component with the** *least number* **of significant figures.**

$$\frac{\overbrace{541}^{3} \times \overbrace{2.067}^{4}}{\underbrace{39}_{2}} = 28.673$$

*2* ≡ **least number!**

*we round off here*

**Result: 29**

# A comprehensive example from the Chemistry lab

- **We want to prepare 500 mL of a standard solution of $KMnO_4$ ~ 0.2 M.**
  - **Let us calculate roughly the amount of $KMnO_4$ needed to have approximately a 0.2 M solution:**

$$m_{KMnO4} = M_{KMnO4} \times \text{Molarity} \times V_{soln}$$
$$\cong 158\ (g/mol) \times 0.2\ (mol/L) \times 0.5\ (L)$$
$$\cong 16\ g$$

  - **Now weigh accurately about 16 g of $KMnO_4$ on a PE 160 Mettler balance:**

# Weighing of $KMnO_4$

| | |
|---|---|
| Mass of beaker | **19.346 g** |
| Mass of beaker + permanganate sample | **35.297 g** |
| Mass of permanganate sample | **15.951 g** |

# Preparation of a primary standard

- **Transfer this accurate amount of $KMnO_4$** quantitatively **into a 500 mL volumetric flask.**
- **Make up to the mark. The uncertainty on the volumetric flask is marked ± 0.25 mL. Thus the volume is reported as 500.0 mL** ≡ 0.5000 L **(4 Sig. Fig.).**

*Quantitative transfer operation*

## Calculation of the molarity of the primary standard

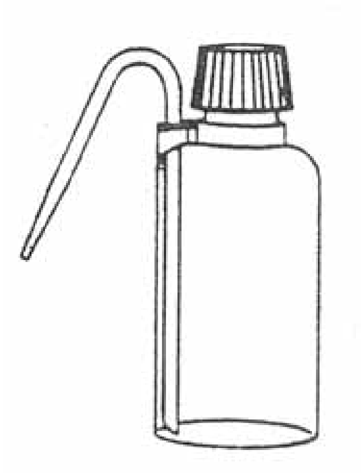
– **Now, I can calculate the accurate** Molarity **of my** standard solution:

$$\text{Molarity} = \frac{m_{KMnO_4}}{M_{KMnO_4} \times V_{soln}}$$

$$= \frac{15.951\ (g)}{158.04\ (g/mol)\ \times 0.5000\ (L)}$$

$$= 0.201860288\ \ M$$

$$\rightarrow \underline{0.2019\ M}\ \ (4\,\text{sig. fig.})$$



*Wash bottle*

## Another illustrative calculation

$$0.876 \equiv 0.9$$

$$= \frac{(4.1\ -\ 3.224)\ \times\ 7.05}{0.00698}$$

$$= 909.025788$$

$$= 9 \times 10^2$$