Chapter 1

Scientific Method, Units and Density

Sections 1.2-1.8

Scientific Method

- **Observation** of a phenomenon
- **Hypothesis** that attempts to explain the observation
- **Experiment** or study which will provide an answer to that hypothesis
- The results of that experiment either leads to a new hypothesis because of a failure, or a law and subsequently a theory which explains the observation.
- A theory can be modified constantly based on further predictions and corresponding experiments until it fully explains the observed phenomenon.

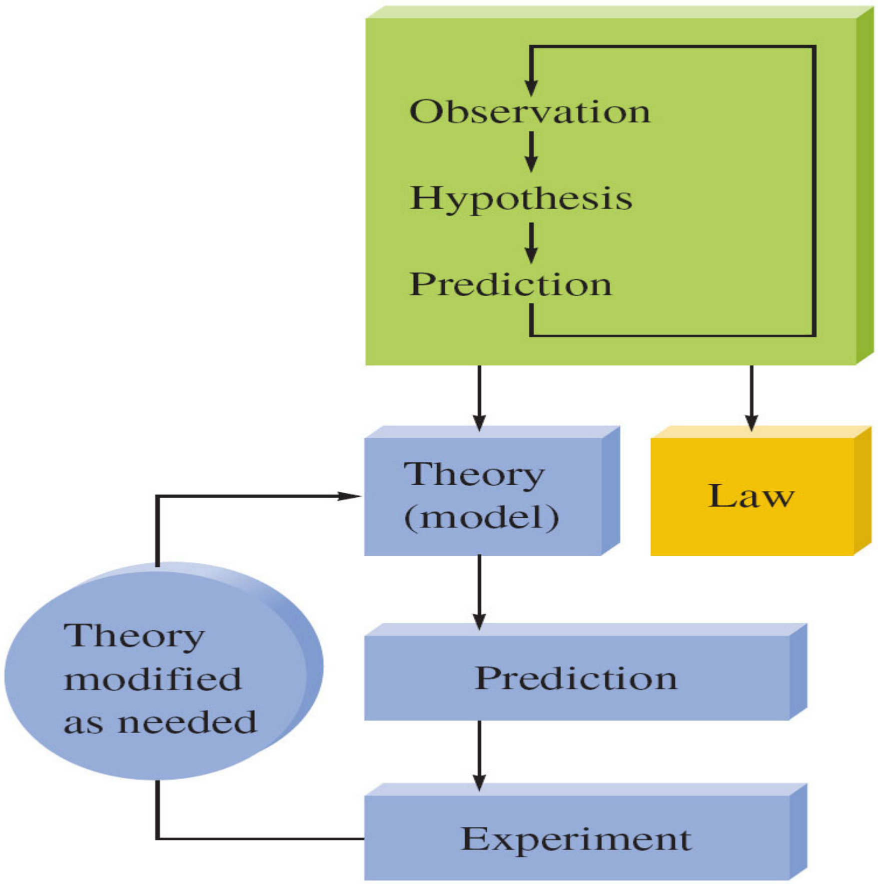
Scientific Method Example

- Observation: Galileo clumsily drops many coins on the ground and wonders whether weight affects the time taken for an object to reach the ground.
- Hypothesis: A heavier object falls faster than a lighter object because of its larger mass.
- Experiment: He drops two objects of known mass (e.g 5kg and 10kg), of the same size, from the tower of Pisa and uses an assistant to help measure the time it takes for each object to reach the ground.

Findings

- Galileo found that both objects land at the same time regardless of mass.
- This is what is known as a **law**.
- It will take many more years before the phenomenon is satisfactorily explained using gravitational **theory**.

Scientific Method



- A **law** tells you what happens; it does not explain why the phenomenon occurs.
- A **theory** explains why the law exists, and is improved and refined over time through further experiments.

SI: International System for Units

- Metric System:

Table 1.1 ▶ The Fundamental SI Units

Physical Quantity	Name of Unit	Abbreviation
Mass	kilogram	kg
Length	meter	m
Time	second	s
Temperature	kelvin	K
Electric current	ampere	A
Amount of substance	mole	mol
Luminous intensity	candela	cd

  - Note the abbreviation for seconds.

SI Prefixes

Table 1.2 ▶ The Prefixes Used in the SI System (Those most commonly encountered are shown in blue.)

Prefix	Symbol	Meaning	Exponential Notation*
exa	E	1,000,000,000,000,000,000	10^{18}
peta	P	1,000,000,000,000,000	10^{15}
tera	T	1,000,000,000,000	10^{12}
giga	G	1,000,000,000	10^{9}
mega	M	1,000,000	10^{6}
kilo	k	1,000	10^{3}
hecto	h	100	10^{2}
deka	da	10	10^{1}
—	—	1	10^{0}
deci	d	0.1	10^{-1}
centi	c	0.01	10^{-2}
milli	m	0.001	10^{-3}
micro	μ	0.000001	10^{-6}
nano	n	0.000000001	10^{-9}
pico	p	0.000000000001	10^{-12}
femto	f	0.000000000000001	10^{-15}
atto	a	0.000000000000000001	10^{-18}

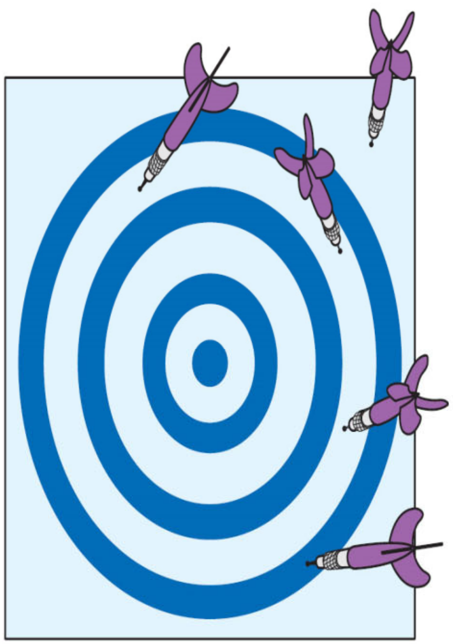
- You should know prefixes and symbols from nano to mega

Precision and Accuracy

- Precision: Degree of agreement between several measurements of the same quantity.
- Accuracy: Agreement of a particular value with the true value.

A game of darts

- Random	Precise	Accurate Precise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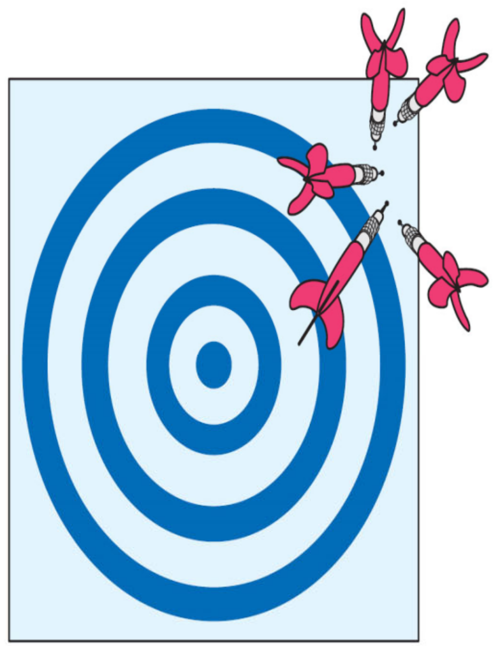
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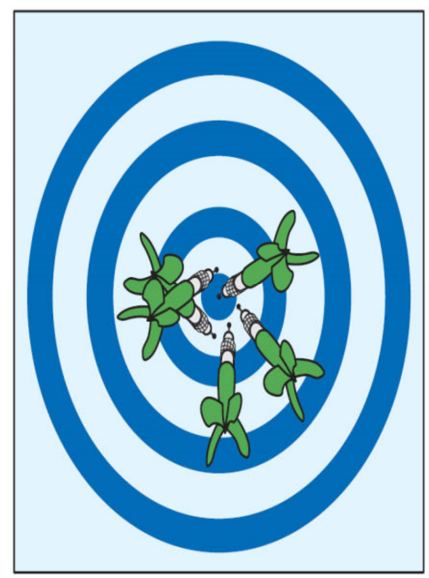
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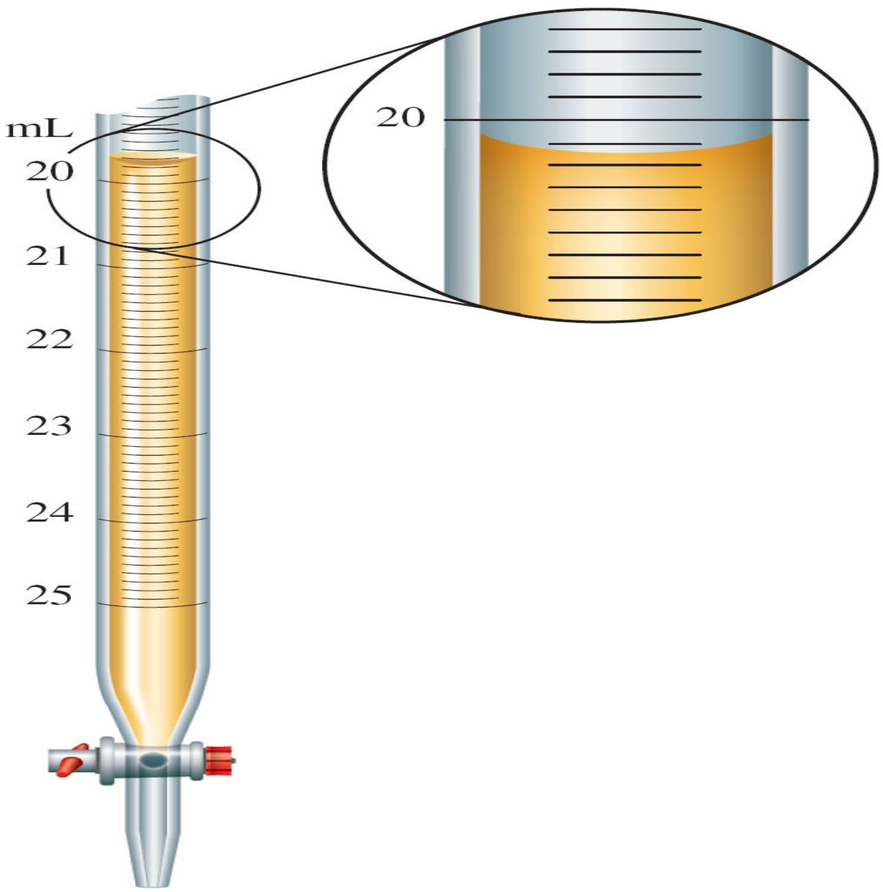
Uncertainty



- Does this buret read 20.14 mL, 20.15 mL or 20.16 mL in the larger image?
- We can say that the first three digits are certain but that the last digit (one hundredth) is uncertain.
- All chemical glassware in general will have an uncertainty

Uncertainties in the lab

- Various instruments in the lab will have various uncertainties, the number of decimal places you **must** report will depend on the instrument used.
- For example, if one is measuring for volume,
  - Beaker, Erlenmeyer: No decimal place reported
  - Graduated Cylinder: One decimal place reported
  - Buret, Pipette, Volumetric Flask: Two decimal places reported
  - An electronic device (such as a balance): As many decimal places as the device reports

Significant Figures

- When doing **laboratory related work**, you must report any values **measured** with the right amount of decimal places.
- In addition, if a number is obtained from **calculations** on values that were measured, significant figures (S.F.) rules must be applied.

Logic behind SF

- The logic behind significant figures is that you can't report something that you don't know.
- If you report 2.0000g, you are saying that you are certain about all these zeros. If you are not, then you need to report 2g.
- These rules are followed stringently by chemists in many industries such as the food industry where food labels need to have be correct or risk lawsuits.

Counting Significant Figures

- Balance A reads 25.4g
- Balance B reads 25.397g

Measurement A has 3 significant figures (SF)

Measurement B has 5 SF

1) All non zero digits are counted as significant figures.

Zeros and their significance

2) Zeros in between non zero digits always have significance.

- Example: 403 has 3 SF.

3) Leading zeros have no significance

- Example: 0.00029 and 00043 have 2 SF.

4) Trailing zeros at the end of a number are significant only if the number contains a decimal point.

- Example: 0.03000,15670000, 4300. have 4 SF.

Examples

Give the right number of SF in the following:

- 2005
- 2460
- 2460.
- 1.350
- 002.204

Exponential Notation

- Exponential notation is a way to represent a number, and has the following form:
  - A number between one and ten multiplied by 10 to the power of an exponent.
- The number 0.0000092 can be written in exponential notation as 9.2×10^{-6}
- Similarly, instead of writing 92000000 it is more convenient to write it as 9.2×10^{7}

Exponential Notation

- If one is following significant figures rules (laboratory), the number of significant figures from the conversion must remain the same.
- Example: A balance reads 124.150g. Convert this to the exponential notation while respecting the number of significant figures.

SF: Addition and Subtraction

- When adding or subtracting two or more numbers, all digits are kept until the final stage. The final answer is then expressed in the **least number of decimals**. Number of significant figures is irrelevant!!
- For example:

1.0+1.11+1.139 = 3.249 which becomes 3.2

Note about rounding

- We will use simple rounding rules for this course.
  - If the last digit is equal to or greater than 5, round up.
  - If the last digit is smaller than 5, round down.

Examples

- $50+250 =$

- $2.2-2.15+4000=$

Multiplication and Division

- When performing multiplication and division, the answer is rounded off to the measurement with the **least number of significant figures**
- For example:
  - 36 x 208 = 7488 which becomes 7500 or $7.5\text{x}10^3$

Examples

- $50 \times 2.0 \times 1.000$

- $$\frac{12.2 + 5.80}{4.5 \times 2.0}$$

Exact Numbers

- Some numbers are exact, because of counting (such as when doing an average), or because it is a known value; for example most unit conversion factors are considered exact.

  (e.g 1 in = 2.54 cm)
- They have an infinite number of significant figures and won't be considered when rounding the final answer.

  For example: A nickel weighs exactly 5.00g. What is the weight of two of these nickels?

Answer: 2 x 5.00g = 10.0g (The 2 is an exact number and thus we express our answer with 3 significant figures)

Visualizing SF

$$\frac{\overset{3SF}{12.2} \times \overset{2\,SF}{3.1} + 3.39}{4.23} = \frac{\overset{2\,SF!}{\underline{37}.82} + 3.39}{4.23} \Rightarrow \frac{\overset{0\,dec}{38} + \overset{2\,dec}{3.39}}{4.23}$$

$$= \frac{\overset{0\,dec!}{41.21}}{4.23} \Rightarrow \frac{\overset{2\,SF}{41}}{\underset{3SF}{4.23}} = \overset{2\,SF!}{\underline{9.7}4231} \Rightarrow 9.7$$

Final Answer

Calculator Answer

Note about SF

- For all **laboratory** related activities, you need to keep track of significant figures.
- For all classroom related activities (in class problems, quizzes, exams and even the final exam), I will not grade significant figures unless it is a question on SF.
  - As a general practice, try to keep as many decimals (use the calculator answer function) during intermediate steps and round up at the end to a reasonable amount.
  - Example: Your calculator gives you 15.2345231 as the answer.
    - Reasonable: 15.2, 15.23
    - Unreasonable: 15.2345231

Dimensional Analysis

- To facilitate conversion between units which we will be doing a lot, we will use dimensional analysis
- It is also sometimes called factor labeling.
- It is a much more efficient version of cross products.

Factor Labelling Method Steps

- 1) List all necessary conversion factors (i.e 1 hour = 60 minutes) as well as any information that is given to you.
- 2) Convert each conversion factor into a fraction while keeping the units. (1 hour/60 minutes or 60 minutes/1hour)
- 3) List the desired unit on the left side of the equation.
- 4) To obtain that unit, you must place the conversion factor fraction or your initial information the "right way" to get your desired unit.
  - Example: I want to convert something in hours, I will have to list my conversion factor as 1 hour / 60 minutes on the right side of my equation.
- 5) Multiply your unit by other conversion factors or given information until you obtain the desired unit inverting as necessary.

Example

Light travels at approximately $3x10^8$ m per second. What is its speed in km/h?

Dimensional Analysis

- If the units don’t cancel each other out to give the expected unit (that of the quantity you’re trying to convert to), then the answer is **wrong**.
- This method is extremely useful because you do not have remember formulas as long as you know the conversion factor.

Weight and Mass

- Distinction between the two terms:
  - Your mass is always the same whether you're on Earth, on the moon, or in space. Your weight will change however based on gravity
  - In this course, we will restrict ourselves to chemistry **on Earth**, so mass and weight will be terms used interchangeably.

Temperature Scales

- Kelvin (K): Official SI unit for temperature. Used widely in science; it cannot be a negative value.
- Celsius (°C): Used in science and everyday life.
- Fahrenheit (°F): Used in the United States and in engineering sciences.

Temperature Conversions

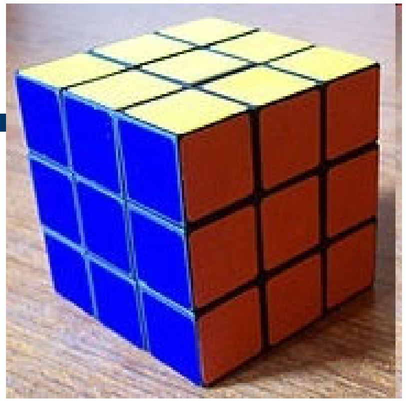
- $T_K = T_C + 273$ (273.15 is the exact value but we will drop the last two decimals)
  - T_K is the temperature in Kelvin and T_C is the temperature in degrees Celsius
- Rearranging yields $T_C = T_K - 273$

For example: Express 20^{o}C in Kelvins.

$T_K = T_C + 273$

$T_K = 20 + 273 = 293 \Rightarrow 293$ K

Volume Measurements



- By definition, 1 dm^3 = 1 L

Alternatively we can also say **1 cm^3 = 1 mL**

Proof:

$$1\ L = 1\ dm \times 1\ dm \times 1\ dm = 1\ dm^3$$
$$= 10\ cm \times 10\ cm \times 10\ cm = 1000\ cm^3$$
$$1000\ mL = 1000\ cm^3 \quad (1\ mL = 1\ cm^3)$$

- A useful way to quicken conversion involving cubes is to think of it as 1 dm^3 as 1 $(dm)^3$.

Example

- How many dm^3 and μm^3 of water are there in 1.24 ML of water?

Density

- Measurement of mass of a substance per unit of volume.
  - Example: g/mL or g/cm^3
- The density of water is approximately 1g/mL or 1kg/L at room temperature
- Density of liquids can be obtained by measuring the weight of a known volume of liquid.
- If two liquids don't mix (such as oil and water), the one with the lower density will float over the other.

Density Problems

- To solve density problems, we will use dimensional analysis.
  - The main thing to remember is that density is expressed in **units of mass/units of volume.**

Example

- A solution of 2.2 L of isopropyl alcohol weighs 1.734 kg. What is the density of isopropyl alcohol in g/cm^3? If isopropyl alcohol is not miscible (does not mix) with water, will it be the top or bottom layer?