Keys to the Study of Chemistry

- **Chemistry** is the study of matter, its properties, changes, and the energy associated with these changes
- **Matter** is everything that has mass an occupies space
  - **Pure substances**
  - **Mixtures**

1.1 Fundamental Definitions

- **Changes of matter**
  - **Physical** – changes in the physical form of matter, but not in its chemical identity (e.g., boiling, melting, mixing, diluting, …)
  - **Chemical** – changes in the chemical identity of matter (e.g., chemical reactions such as rusting of Fe, burning of gasoline, digestion of food, …)

- **Properties of matter**
  - **Physical** – characteristics of matter that can be observed without changing its chemical identity (e.g., mass, density, color, physical state, …)
  - **Chemical** – characteristics of matter related to its chemical change (e.g., hydrogen is a flammable gas that burns in the presence of O_2 to produce H_2O)
- A substance is identified by its own set of physical and chemical properties

- **Physical states of matter**
  - **Solid** – a rigid form of matter with definite volume and shape
  - **Liquid** – a fluid form of matter with definite volume but not shape
  - **Gas** – a fluid form of matter with no definite volume or shape (no surface)
- In general, changes in the physical state are reversible and can be achieved by changing temperature and pressure

- **Macroscopic and microscopic properties and events**
  - **Macroscopic** – observable properties and events of large visible objects
  - **Microscopic** – result from changes at a much smaller (atomic) level not visible by the naked eye
- Macroscopic properties and events occur as a result of microscopic properties and events

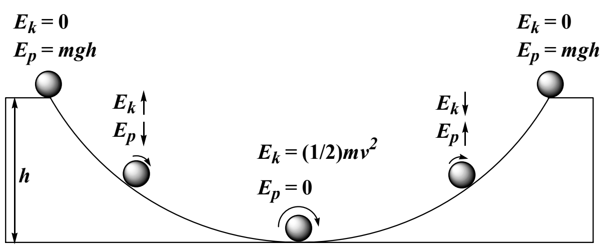
Examples:

- Define the following as physical or chemical properties or changes:
  - A stove becomes red-hot
  - The leafs of a tree turn yellow
  - Lead is a dense metal
  - Acetone is quite volatile (easily vaporized)
  - Iron rusts when exposed to air
  - Gasoline is flammable

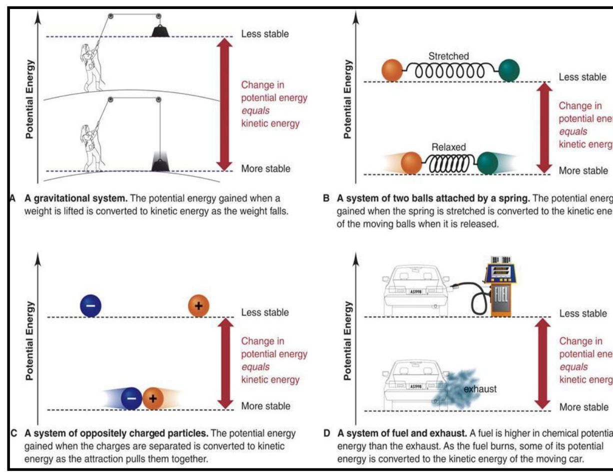
- **Energy** – the ability to do work
  - **Potential energy** – due to position or interaction
  - **Kinetic energy** – due to motion
  - **Total energy** – sum of potential and kinetic energy
- Law of **conservation of energy** – the total energy of an isolated object (or a system of objects) is constant
  - Energy is neither created nor destroyed – it is only converted from one form to another

Conservation of Energy

$$E_{tot} = E_k + E_p = constant$$



Note: The friction is neglected.



A **A gravitational system.** The potential energy gained when a weight is lifted is converted to kinetic energy as the weight falls.

B **A system of two balls attached by a spring.** The potential energy gained when the spring is stretched is converted to the kinetic energy of the moving balls when it is released.

C **A system of oppositely charged particles.** The potential energy gained when the charges are separated is converted to kinetic energy as the attraction pulls them together.

D **A system of fuel and exhaust.** A fuel is higher in chemical potential energy than the exhaust. As the fuel burns, some of its potential energy is converted to the kinetic energy of the moving car.

1.2 The Scientific Method

Observations : Natural phenomena and measured events; universally consistent ones can be stated as a **natural law**.

↓

Hypothesis: Tentative proposal that explains observations or natural laws.

revised if experiments do not support it

↓

Experiment: Procedure to test hypothesis; measures one variable at a time.

↓

Model (Theory): Formal explanation of experimental data or natural laws; predicts related phenomena.

altered if predictions do not support it

↓

Further Experiment: Tests predictions based on model.

1.3 The Unit Conversion Method

- **Units of measurement**
  - Measurements – quantitative observations
  - Units – standards used to compare measurements (yard → standard for comparison of length measurements)
  - A measured quantity is reported as a **number** and a **unit**

(Measured quantity) = number × unit

5.5 seconds = 5.5 × 1 s

- **Units in calculations**
  - Units are treated just like pure numbers

  Area = 4 in × 6 in = (4×6)(in×in) = 24 in²
  - Systems of units (metric, English, SI, ...)
  - Equalities between units

  1 in = 2.54 cm **1 mi = 1.609 km**
- **Conversion factors** – ratios between two equal or equivalent units (derived from equalities)

$$\frac{1\ in}{2.54\ cm} = 1 \quad or \quad \frac{2.54\ cm}{1\ in} = 1$$

- **Unit conversions** (old unit → new unit)
  - Quantity remains the same; units change

$$new\ unit = old\ unit \times (conv.factor)$$

$$conv.factor = \frac{new\ unit}{old\ unit}$$

$$new\ unit = old\ unit \times \frac{new\ unit}{old\ unit}$$

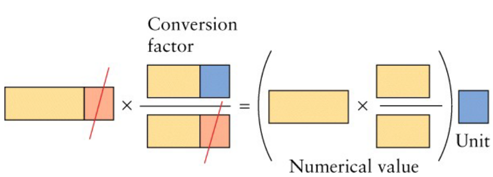
  - The old units cancel

Example: Convert **5.13 inches** to **centimeters**.

old unit → *in* **new unit → *cm***

1 *in* = 2.54 *cm* → conversion factor = [2.54 *cm*/1 *in*]

$$5.13\ in \times \frac{2.54\ cm}{1\ in} = \frac{5.13 \times 2.54}{1} \times \frac{in \times cm}{in} = 13.0\ cm$$



Example:

- The gas mileage of a car is **35 mi/gal**. How many **km** can the car travel on a full **10 gal** tank of gas?

  1 mi = 1.609 km

$$10\ gal \times \frac{35\ mi}{1\ gal} = 350\ mi$$

$$350\ mi \times \frac{1.609\ km}{1\ mi} = 563\ km$$

1.4 Measurement in Scientific Study

- Systems of units (metric, English, SI, ...)
- The International System of units (SI)
  - Based on the metric system
  - SI base units

Table 1.2 SI Base Units

Physical Quantity (Dimension)	Unit Name	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

- Prefixes used with SI units (denote powers of 10)
  - Used to express very small or very large quantities

Table 1.3 Common Decimal Prefixes Used with SI Units

Prefix*	Prefix Symbol	Meaning: Number	Meaning: Word	Multiple†
tera	T	1,000,000,000,000	trillion	10^{12}
giga	G	1,000,000,000	billion	10^{9}
mega	M	1,000,000	million	10^{6}
kilo	k	1,000	thousand	10^{3}
hecto	h	100	hundred	10^{2}
deka	da	10	ten	10^{1}
—	—	1	one	10^{0}
deci	d	0.1	tenth	10^{-1}
centi	c	0.01	hundredth	10^{-2}
milli	m	0.001	thousandth	10^{-3}
micro	μ	0.000001	millionth	10^{-6}
nano	n	0.000000001	billionth	10^{-9}
pico	p	0.000000000001	trillionth	10^{-12}
femto	f	0.000000000000001	quadrillionth	10^{-15}

- **Examples:**

1 mm = 10^{-3} × (1 m) = 10^{-3} m

1 MW = 10^{6} × (1 W) = 10^{6} W

1 µs = 10^{-6} × (1 s) = 10^{-6} s

1 ng = 10^{-9} × (1 g) = 10^{-9} g

- **Mass and weight**
  - Mass is constant (depends on the amount of matter)
  - Weight can vary with the strength of the gravitational field
  - Mechanical balances actually measure mass

Example:

A jet engine consumes **1.1 gal** of fuel per second. How many **liters** of fuel does the engine need in order to operate for **1.5 hours**?

1 gal = 3.785 L **1 h = 60 min = 3600 s**

Plan:

1.1 gal/s → ? L/s

1.5 Hours → ? minutes → ? seconds

Seconds × L/s → ? L

Example (cont.):

$$1.1\,\frac{\text{gal}}{\text{s}} \times \left(\frac{3.785\,L}{1\,\text{gal}}\right) = 4.2\,\frac{L}{\text{s}}$$

$$1.5\,h \times \left(\frac{60\,min}{1\,h}\right) \times \left(\frac{60\,s}{1\,min}\right) = 5400\,s$$

$$5400\,s \times \left(\frac{4.2\,L}{1\,s}\right) = 22000\,L$$

- **Derived units** (derived from the base units)
  - **Volume (V)** → 1 m^3 = (1 m) × (1 m) × (1 m)

1 mL = 1 cm^3 = (1 cm)×(1 cm)×(1 cm) = (10^{-2} m)×(10^{-2} m)×(10^{-2} m) = (10^{-2}×10^{-2}×10^{-2}) m^3 = 10^{-6} m^3

1 L = 1 dm^3 = 10^{-3} m^3 **1 mL= 10^{-3} L**

  - **Density (d)** → mass (m) per unit volume (V) → ($d = m/V$)

unit of d = (1 kg)/(1 m^3) = 1 kg/m^3

  - **Velocity (v)** → distance (l) per unit time (t) → ($v = l/t$)

unit of v = (1 m)/(1 s) = 1 m/s

- **Extensive properties** – depend on sample size (mass, volume, length, ...)
- **Intensive properties** – independent of sample size (density, temperature, color, ...)

Example:

What is the density of an alloy in g/cm^3, if **55 g** of it displace **9.1 mL** of water?

***$d = m/V$* = (55 g)/(9.1 mL) = 6.0 g/mL = 6.0 g/cm^3**

Example:

- Convert the density of gold, **19.3 g/cm^3**, to **kg/m^3**.

⇒ need to convert both the numerator and denominator **g → kg** **and** **cm^3 → m^3**

1 kg = 10^3 g

1 cm = 10^{-2} m ⇒ 1 cm^3 = $(10^{-2})^3$ m^3 = 10^{-6} m^3

$$d = 19.3\,\frac{g}{cm^3} \times \left(\frac{1\,kg}{10^3\,g}\right) \times \left(\frac{1\,cm^3}{10^{-6}\,m^3}\right) = 19.3\times10^3\,\frac{kg}{m^3}$$

Example:

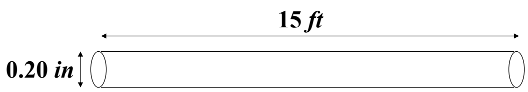
What is the mass in ***kg*** of a **15 *ft*** wire made of an alloy with ***d* = 6.0 g/cm³** if the diameter of the wire is **0.20 *in***?

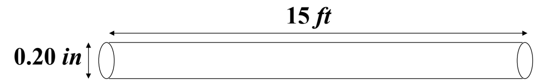
Plan:

Diameter→radius (cm)→cross-section area (cm^2)

Length (cm) × cross-section area → volume (cm^3)

Volume & density → mass (g) → mass (kg)





Radius → $r = 0.20\ in / 2 = 0.10\ in$

$$r = 0.10\ \cancel{in} \times \left(\frac{2.54\ cm}{1\ \cancel{in}}\right) = 0.254\ cm$$

$$A = \pi r^2 = 3.14 \times (0.254\ cm)^2 = 0.203\ cm^2$$

$$l = 15\ \cancel{ft} \times \left(\frac{12\ \cancel{in}}{1\ \cancel{ft}}\right) \times \left(\frac{2.54\ cm}{1\ \cancel{in}}\right) = 457\ cm$$

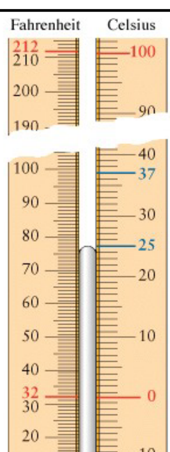
$$V = l \times A = 457\ cm \times 0.203\ cm^2 = 92.7\ cm^3$$

$$m = 92.7\ \cancel{cm^3} \times \left(\frac{6.0\ \cancel{g}}{1\ \cancel{cm^3}}\right) \times \left(\frac{1\ kg}{1000\ \cancel{g}}\right) = 0.56\ kg$$

- **Temperature (*T*)** – a measure of how hot or cold an object is relative to other objects
  - *T* reflects the thermal energy of the object
  - *T* is an intensive property
- **Heat** – the flow of thermal energy between objects
  - Heat flows from objects with higher *T* to objects with lower *T*
  - Heat is an extensive property
  - Heat and temperature are different
- Thermometers
  - Used to measure *T*

- The **Celsius** scale
  - 0ºC → freezing point of water
  - 100ºC → boiling point of water
- The **Fahrenheit** scale
  - 0ºF → freezing point of salt/water mixture
  - 100ºF → body temperature
  - water freezes at 32ºF and boils at 212ºF

⇒100 Celsius degrees ↔ 180 Fahrenheit degrees



$$\left(\frac{180°F}{100°C}\right) = \left(\frac{9°F}{5°C}\right)$$

$$T°F = \left(\frac{9°F}{5°C}\right) T°C + 32°F$$

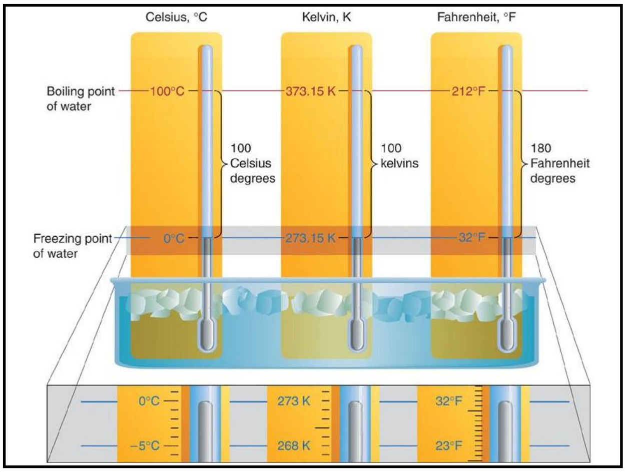
$$T°C = \left(\frac{5°C}{9°F}\right)(T°F - 32°F)$$

- The **Kelvin** scale - absolute temperature scale
  - 0 K → lowest possible temperature
  - 0 K = -273.15°C
  - same size of degree unit as Celsius

⇒water freezes at 273.15 K and boils at 373.15 K

- ***T* K = *T*°C + 273.15**

- ***T*°C = *T* K - 273.15**



Example:

- Convert -40°F in °C and K.
- ***T*°C = (5°C/9°F)×[-40°F - 32°F] =**
  = (5/9)×(-72)°C = -40°C
- ***T* K = -40°C + 273.15 = 233 K**

1.5 Uncertainty of Measurements

- Represents the reliability of measurements
- Reported as: **number ± uncertainty** (4.88 ± 0.05 kg)
- If not reported: assume **±1 in the last reported digit** (3.7 cm → 3.7 ± 0.1 cm)
- Exact numbers – no uncertainty (5 tables, 10 apples, 1 min = 60 s, 1 in = 2.54 cm)

- **Significant figures** – digits of a number known with some degree of certainty
  - All non-zero digits
  - All zeros after the first non-zero digit
  - Exception – trailing zeros in numbers without decimal point are not significant
- More significant figures ↔ less uncertainty

  Examples:

  1.32 → 3 sf

  0.005030 → 4 sf

  4500 → 2 sf

  4500. → 4 sf

- Scientific notation – representation in the form → $A \times 10^{a}$
  - A → a decimal number between 1 and 10
  - a → a positive or negative integer
- **Examples:**

  0.00134 = 1.34×10^{-3}

  134 = 1.34×10^{2}

  - all digits in A are significant

- Examples of significant figures

Decimal notation	Scientific notation	Number of sf
0.751	7.51×10^{-1}	3
0.007 51	7.51×10^{-3}	3
0.070 51	7.051×10^{-2}	4
0.750 100	$7.501\,00 \times 10^{-1}$	6
7.5010	7.5010	5
7501	7.501×10^{2}	4
7500	7.5×10^{3}	2*
7500.	7.500×10^{3}	4

- **Significant figures in calculations**
  - Rounding off (only at the end of a calculation)
    - round up, if next digit is **above 5**
    - round down, if next digit is **below 5**
    - round to the nearest even number, if next digit is **equal to 5** and it is the last nonzero digit of the number (if 5 is not the last nonzero digit, round up)

Examples: Round to 3 sf.

3.7643 → 3.76

3.7683 → 3.77

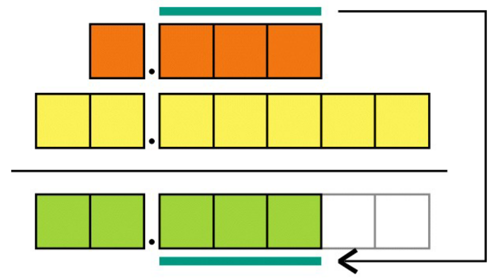
3.7653 → 3.77

3.765 → 3.76

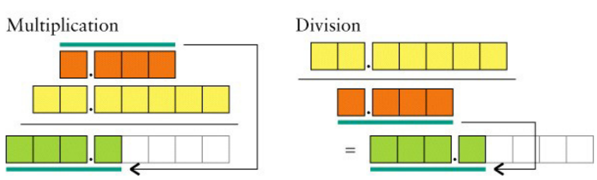
3.765 → 3.76

3.755 → 3.76

- Addition and subtraction
  - the number of decimal places in the result is the same as the smallest number of decimal places in the data



- Multiplication and division
  - the number of significant figures in the result is the same as the smallest number of significant figures in the data



Examples:

0.0354 + 12.1 = 12.1 ← (12.1354)

5.7×0.0651 = 0.37 ← (0.37107)

5.7/0.0651 = 88 ← (87.55760369)

3.568 in × (2.54 cm/1 in) = 9.063 cm

- **Precision and accuracy**
  - Two aspects of uncertainty
- **Precision** – agreement among repeated measurements
  - **Random error** – deviation from the average in a series of repeated measurements (some values higher, some values lower than the average)

small random error ↔ high precision

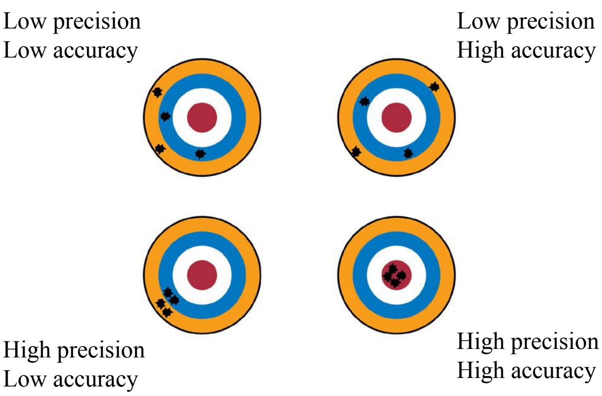
high precision ↔ more sf in the result

- **Accuracy** – agreement of a measurement with the true or accepted value
  - **Systematic error** – deviation of the average from the true value (present in the whole set of measurements – either all high or all low)

small systematic error ↔ high accuracy

- **Instrument calibration** – comparison with a known standard
  - Essential for avoiding systematic error

- Examples of precision and accuracy



Example:

- A car is moving at exactly **60 mi/hr**. Compare the precision and accuracy of the following two series of speed measurements using two different radars.

  A → 61.5, 58.3, 62.7, 63.5, 57.1 (average 60.6)

  B → 62.0, 62.5, 61.8, 62.2, 62.1 (average 62.1)

  A → more accurate, less precise

  B → less accurate, more precise