LECTURER-1

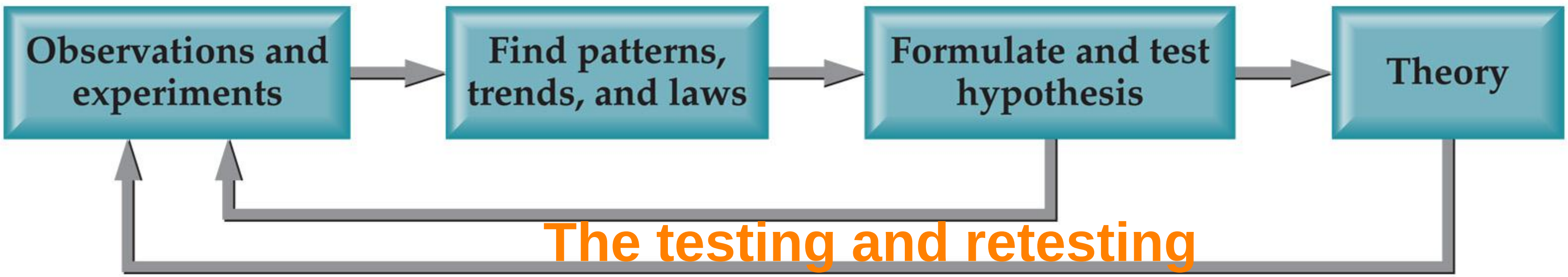
MATTER AND MEASUREMENT

Scientific Method:

A systematic approach to solving problems.

Empirical
Facts

A model

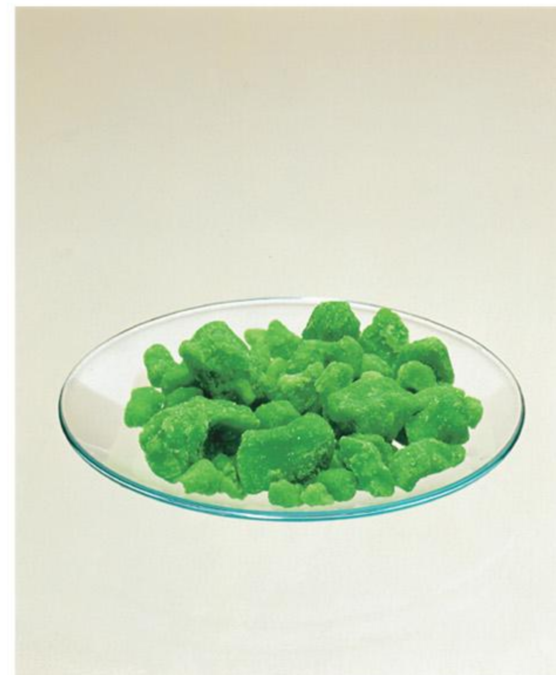


The testing and retesting

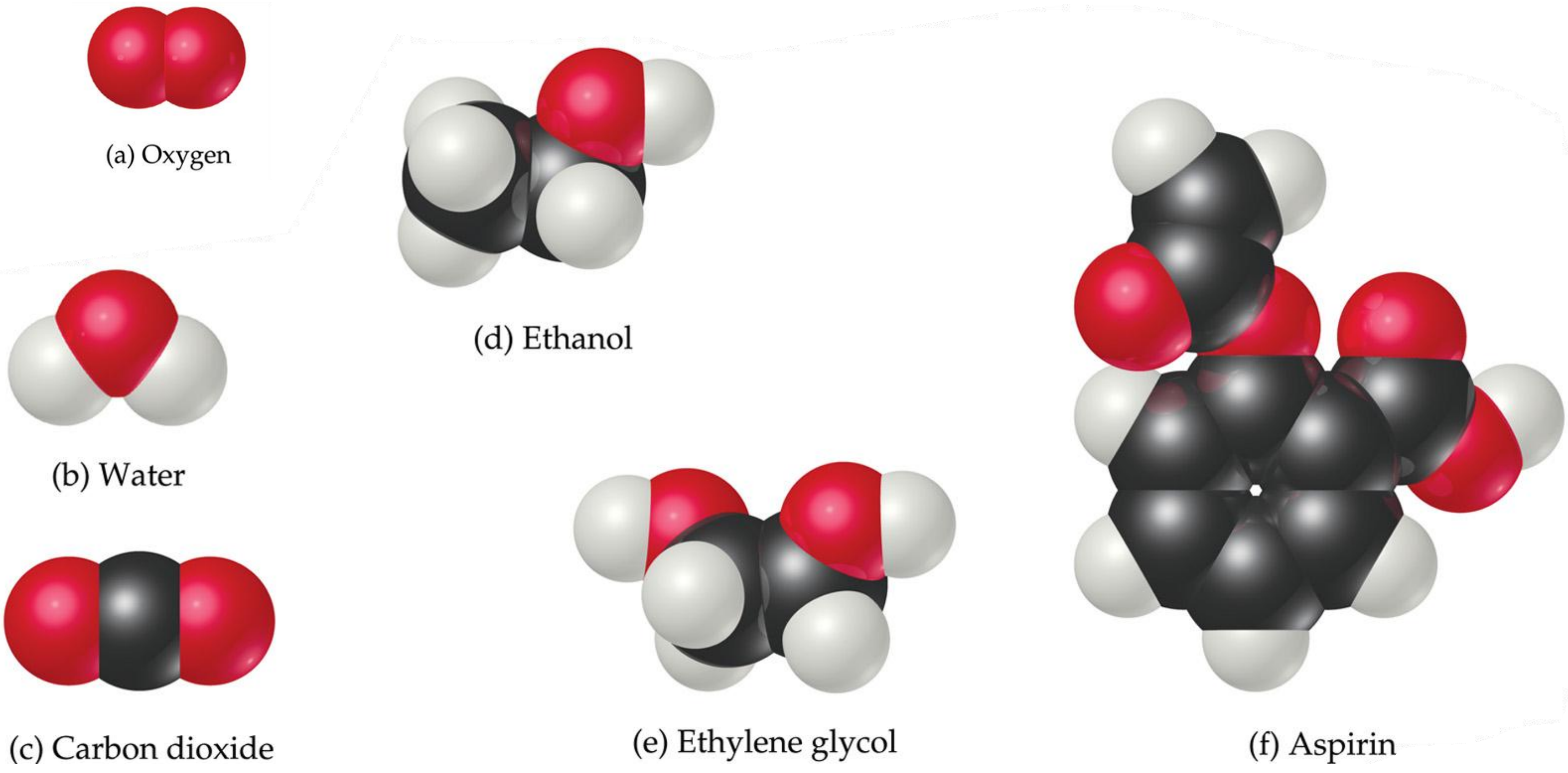
This is what makes it Science!

Matter:

Anything that has mass and takes up space.

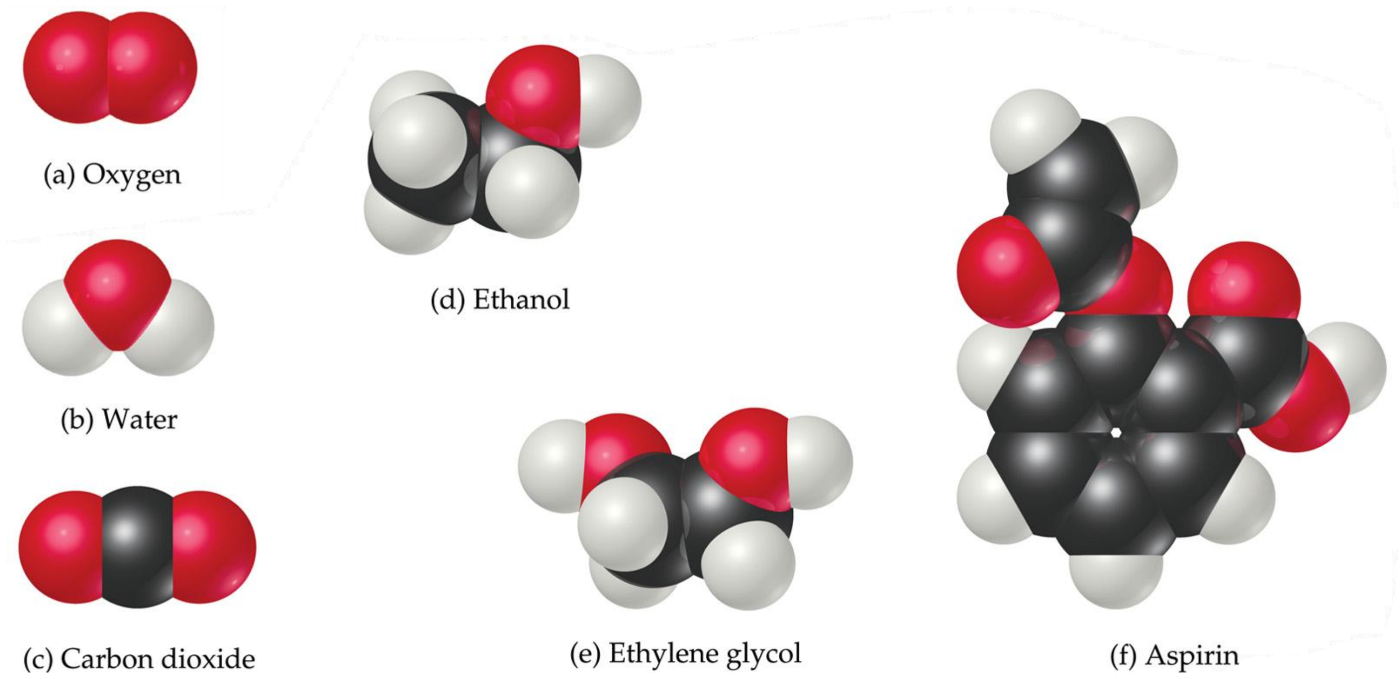


Matter



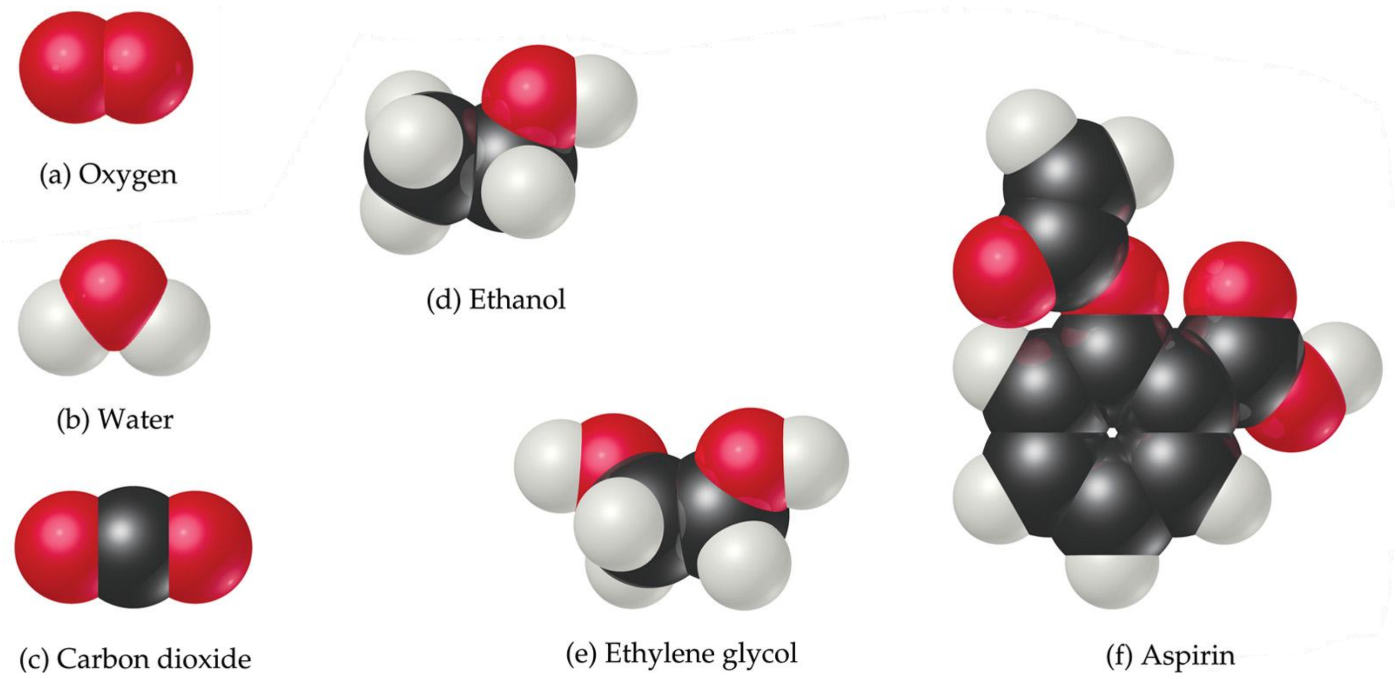
- Atoms are the building blocks of matter.

Matter



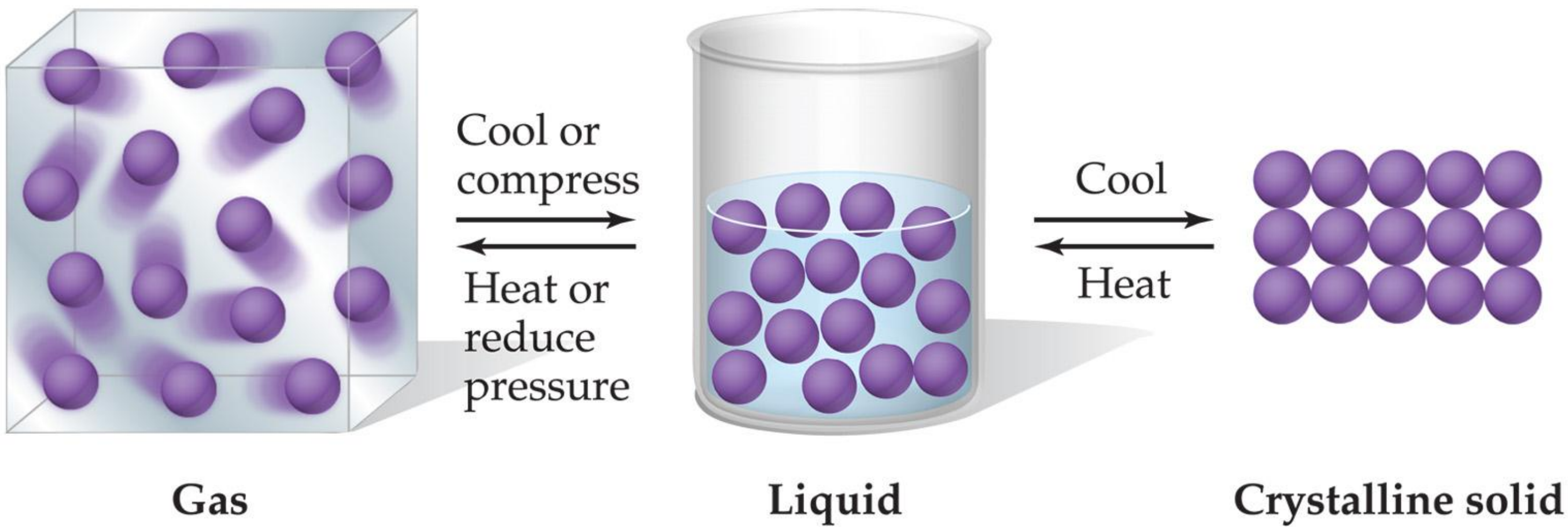
- Each element is made of the same kind of atom.

Matter

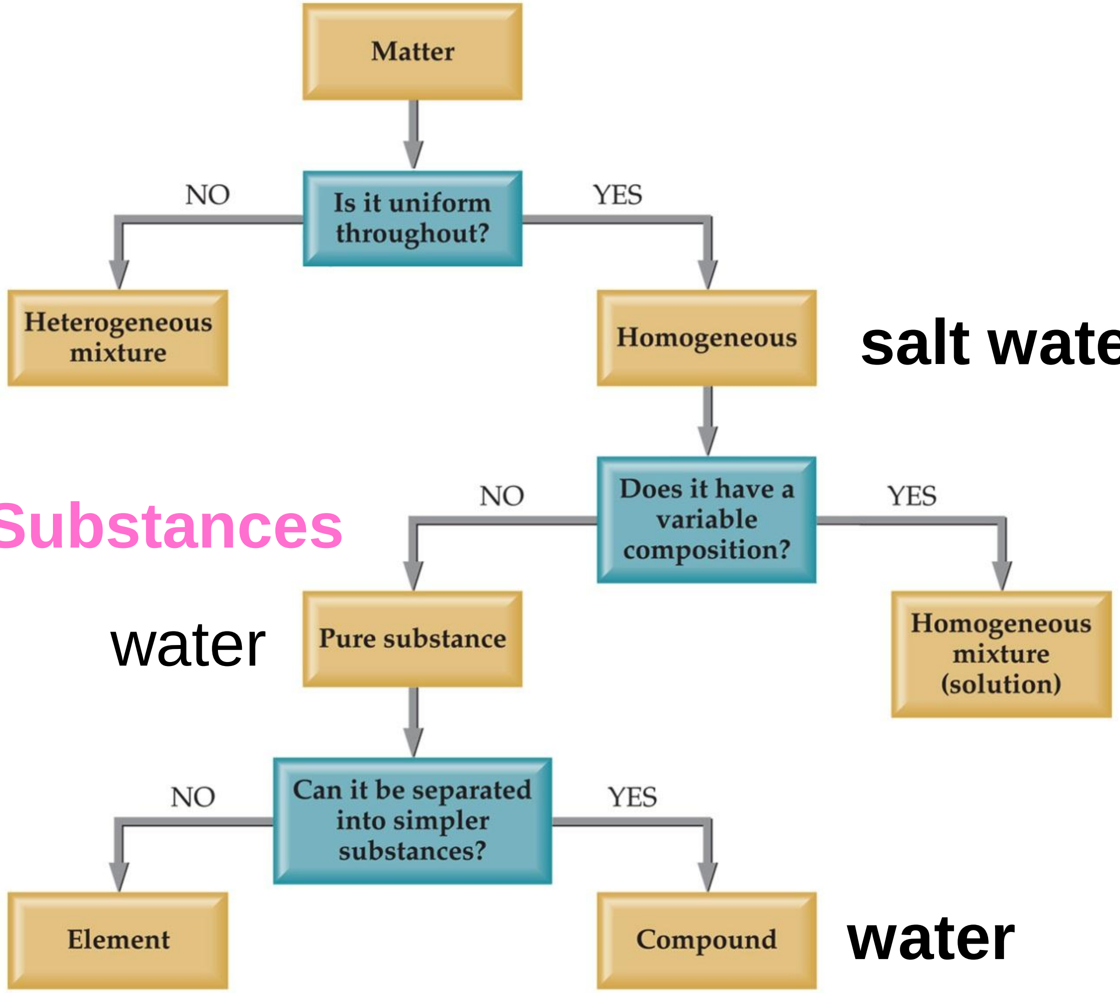


- A compound is made of two or more different kinds of elements.

States of Matter



Classification of Matter



Mud

salt water

Substances

water

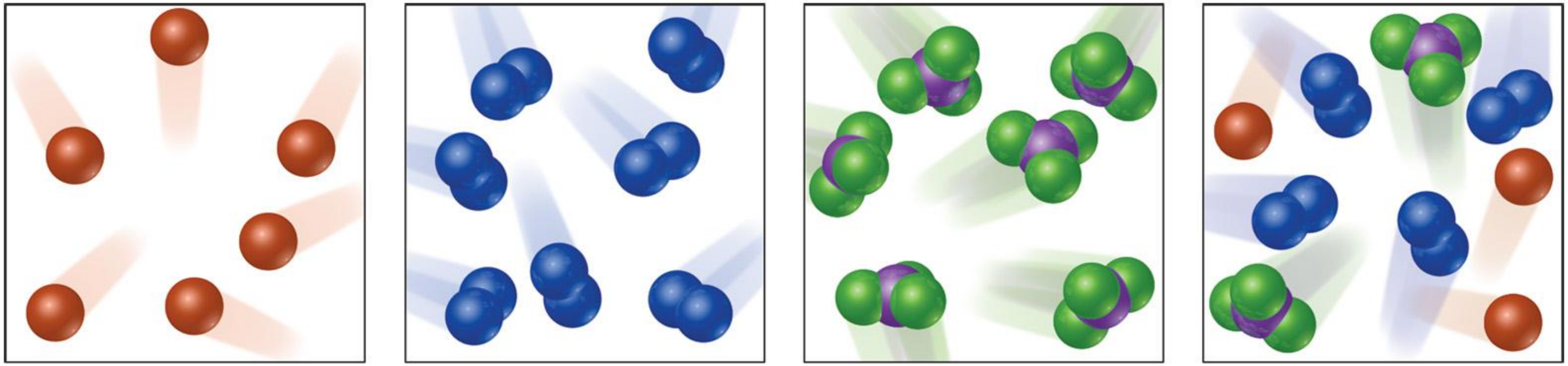
salt water

oxygen

water

Mixtures and Compounds

Element (atoms) | Element (molecules) | Compound (molecules) | Mixture



(a) Atoms of an element

(b) Molecules of an element

(c) Molecules of a compound

(d) Mixture of elements and a compound

He, Ne | N_2, O_2, Cl_2 | CO_2, H_2O, NH_3 | Mix

Properties and Changes of Matter

Properties of Matter

- Physical Properties:
  - □ Must be observed without changing a compound/element into another compound/element.
    - **Boiling point, density, mass, volume, etc.**
- Chemical Properties:
  - □ Can *only* be observed when a compound/element is changed into another compound/element.
    - **Flammability, corrosiveness, reactivity with acid, etc.**

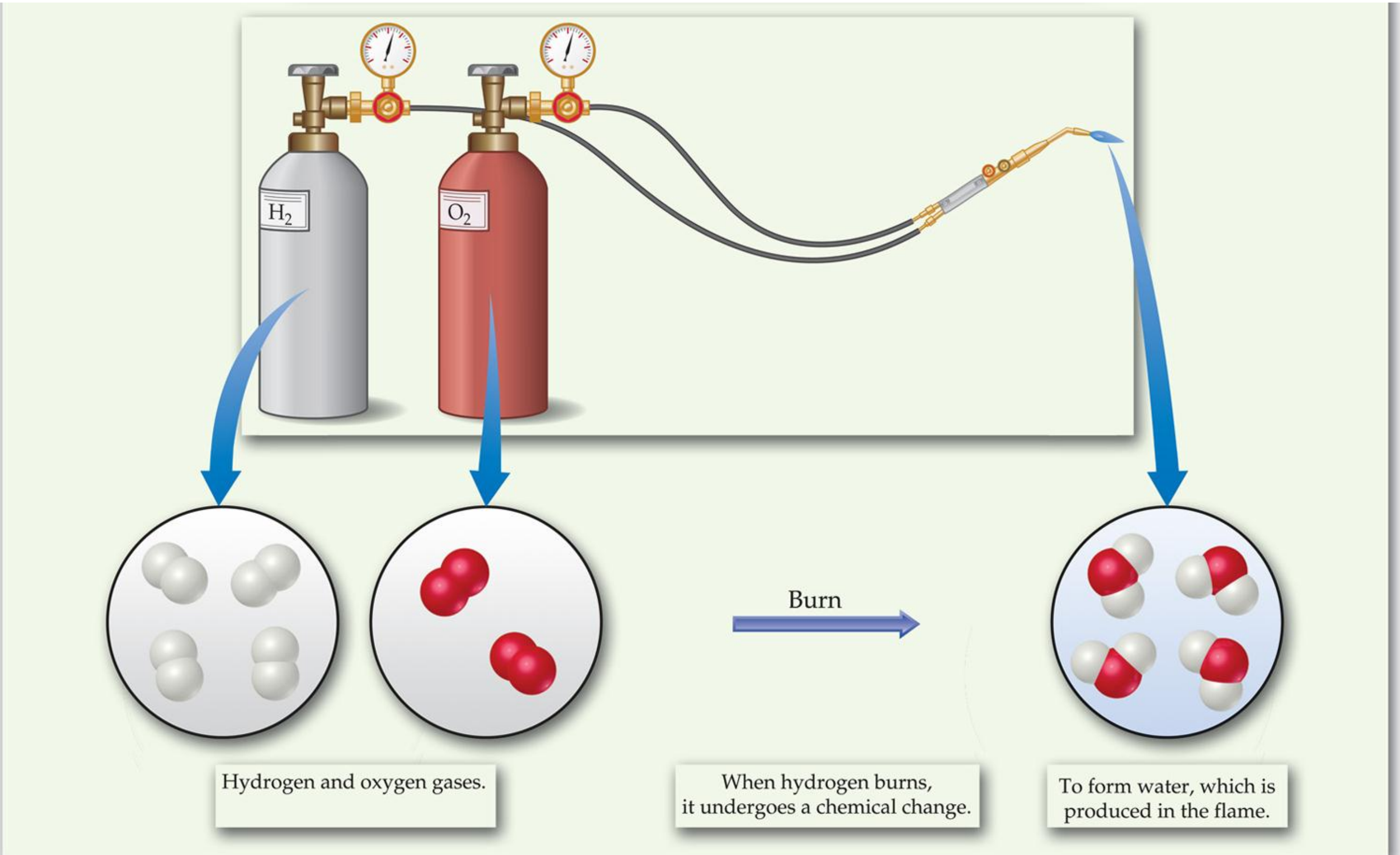
Properties of Matter

- Intensive Properties:
  - □ Independent of the amount of the matter that is present.
    - **Density, boiling point, color, etc.**
- Extensive Properties:
  - □ Dependent upon the amount of the matter present.
    - **Mass, volume, energy, etc.**

Changes of Matter

- Physical Changes:
  - □ Changes in matter that do not change the composition of a substance.
    - **Changes of state, temperature, volume, etc.**
- Chemical Changes:
  - □ Changes that result in new substances.
    - **Combustion, oxidation, decomposition, etc.**

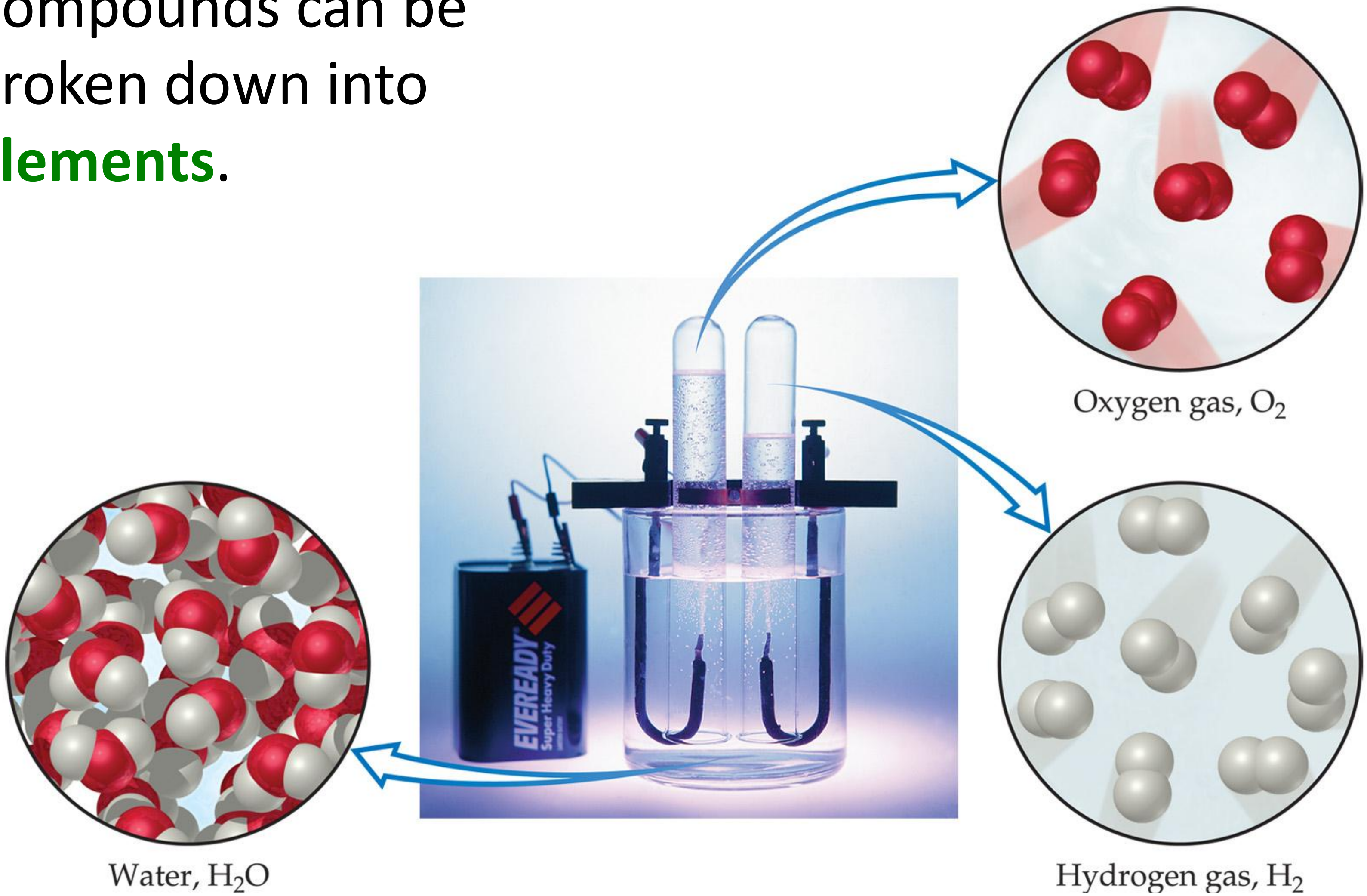
Chemical Reactions



In the course of a chemical reaction, the reacting substances are converted to new substances.

Compounds

Compounds can be broken down into **elements**.



Relative abundance of elements

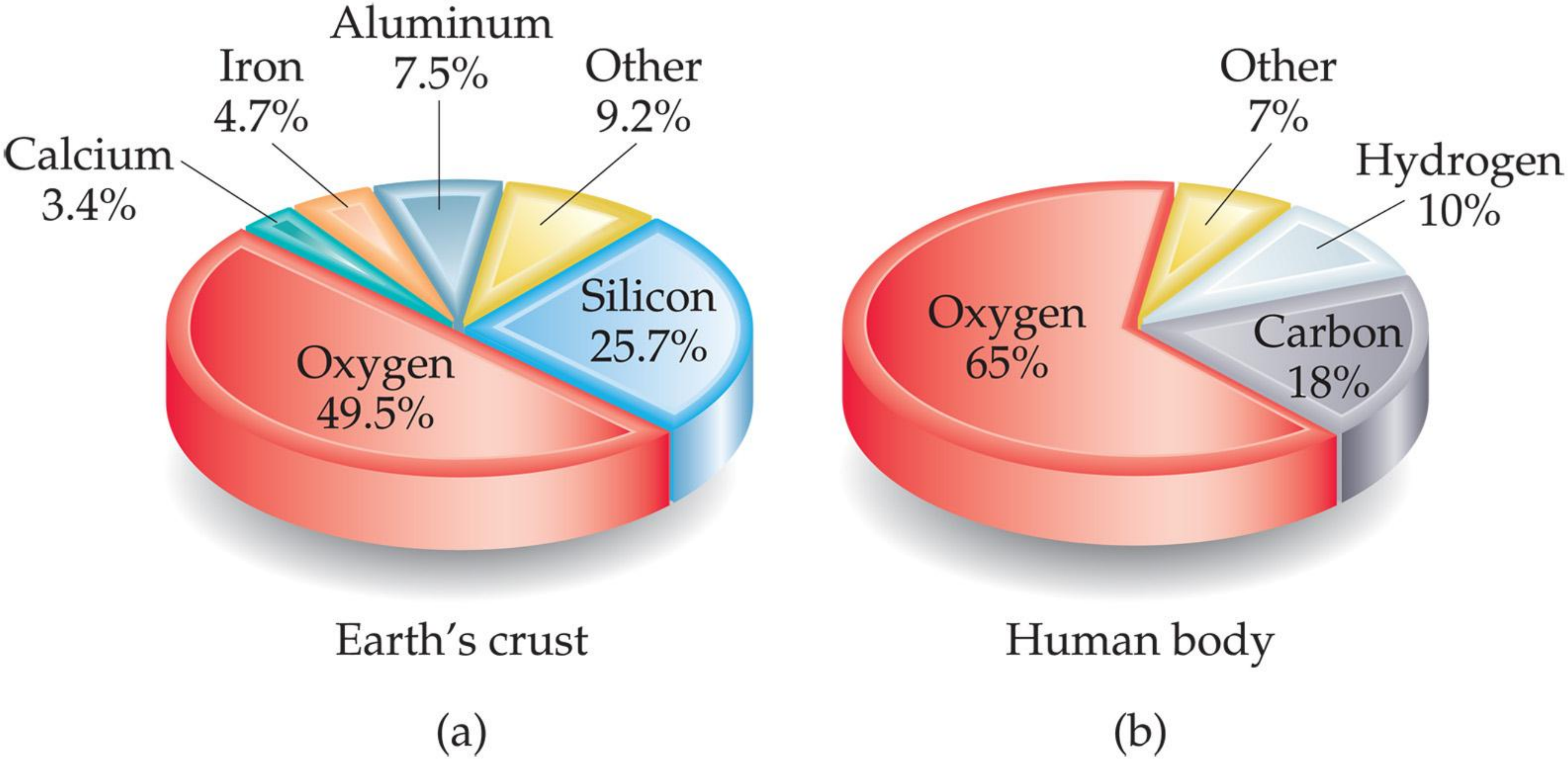


TABLE 1.1 The Top Ten Chemicals Produced by the Chemical Industry in 2002[a]

Rank	Chemical	Formula	2002 Production (billions of pounds)	Principal End Uses
1	Sulfuric acid	H_2SO_4	81	Fertilizers, chemical manufacturing
2	Nitrogen	N_2	73	Fertilizers
3	Oxygen	O_2	53	Steel, welding
4	Ethylene	C_2H_4	52	Plastics, antifreeze
5	Lime	CaO	38	Paper, cement, steel
6	Propylene	C_3H_6	32	Plastics
7	Ammonia	NH_3	29	Fertilizers
8	Chlorine	Cl_2	25	Bleaches, plastics, water purification
9	Phosphoric acid	H_3PO_4	24	Fertilizers
10	Sodium hydroxide	NaOH	20	Aluminum production, soap

[a]Most data from *Chemical and Engineering News*, July 7, 2003, pp. 53, 56.

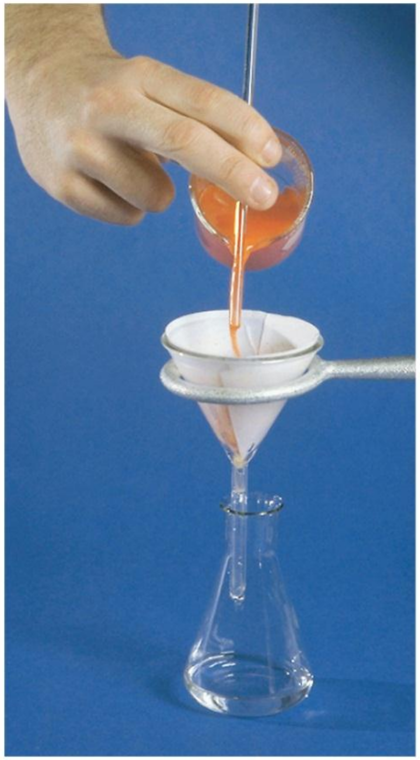
Acids

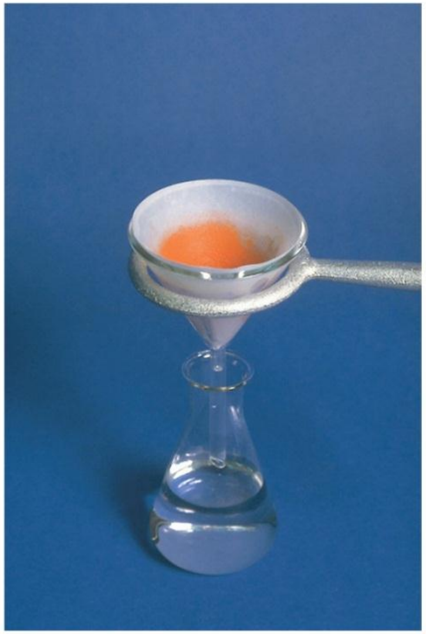
Bases

Pure elements

Separation of Mixtures

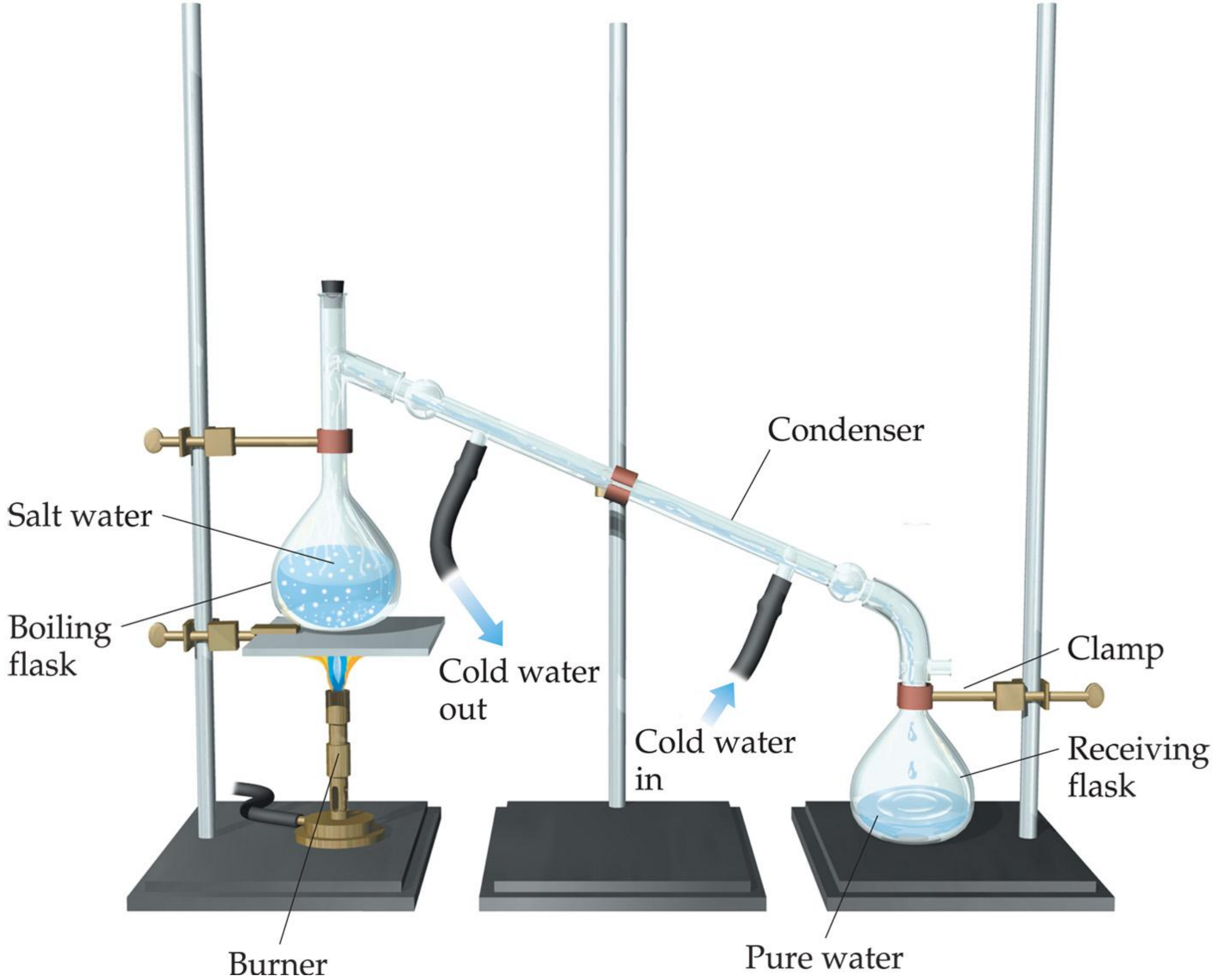
Filtration:





Separates heterogeneous mixture, solid substances from liquids and solutions.

Distillation:



Separates homogeneous mixture of liquids on the basis of differences in boiling point.

Chromatography:

Separates homogeneous mixtures on the basis of differences in solubility in a solvent, or in binding to a solid matrix.

Separation techniques were critical to the development of the basic theories of chemistry.
How do we know there are homogeneous mixtures?
We can separate them.

Units of Measurement

SI Units
Learn! symbols and all!

Physical Quantity	Name of Unit	Abbreviation
Mass	Kilogram	kg
Length	Meter	m
Time	Second	s[a]
Temperature	Kelvin	K
Amount of substance	Mole	mol
Electric current	Ampere	A
Luminous intensity	Candela	cd

[a] The abbreviation sec is frequently used.

- *Système International d'Unités*
- Uses a different base unit for each quantity

Metric System

Prefixes convert the base units into units that are appropriate for the item being measured.

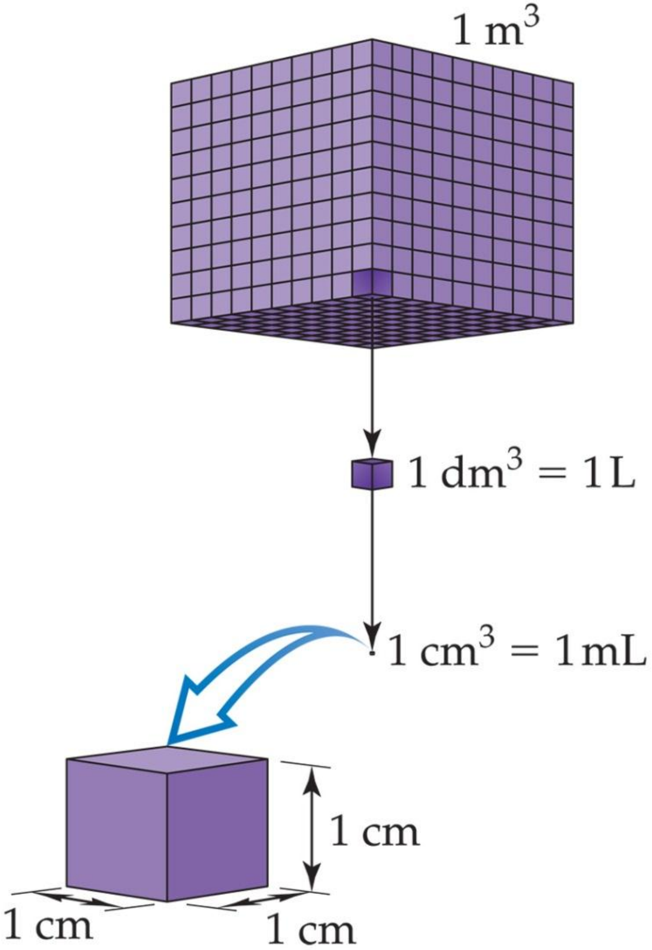
Learn! More important than it looks!!!

Prefix	Abbreviation	Meaning	Example
Giga	G	10^9	1 gigameter (Gm) $= 1 \times 10^9$ m
Mega	M	10^6	1 megameter (Mm) $= 1 \times 10^6$ m
Kilo	k	10^3	1 kilometer (km) $= 1 \times 10^3$ m
Deci	d	10^{-1}	1 decimeter (dm) $= 0.1$ m
Centi	c	10^{-2}	1 centimeter (cm) $= 0.01$ m
Milli	m	10^{-3}	1 millimeter (mm) $= 0.001$ m
Micro	μ[a]	10^{-6}	1 micrometer (μm) $= 1 \times 10^{-6}$ m
Nano	n	10^{-9}	1 nanometer (nm) $= 1 \times 10^{-9}$ m
Pico	p	10^{-12}	1 picometer (pm) $= 1 \times 10^{-12}$ m
Femto	f	10^{-15}	1 femtometer (fm) $= 1 \times 10^{-15}$ m

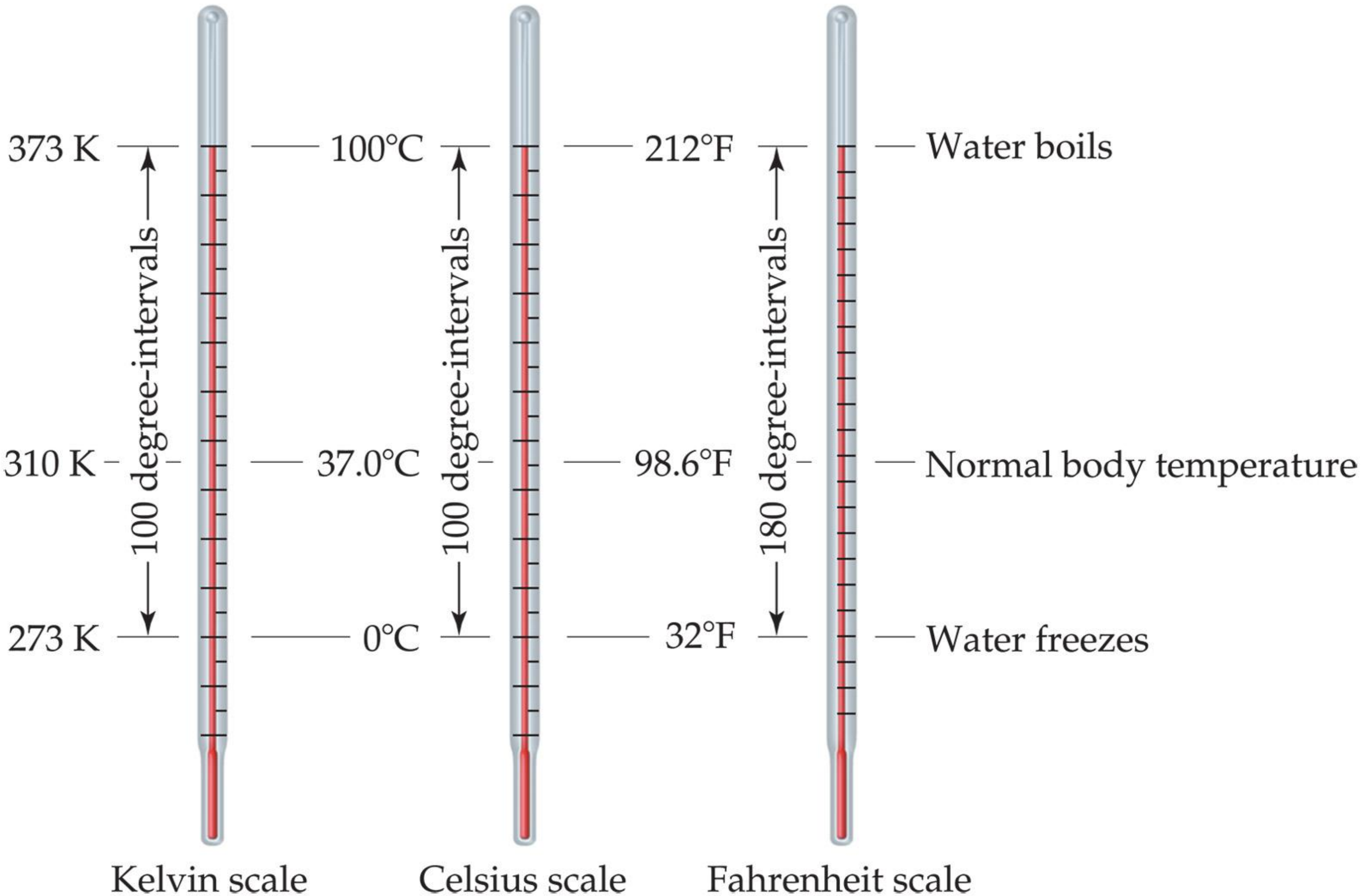
[a] This is the Greek letter mu (pronounced "mew").

Volume

- The most commonly used metric units for volume are the liter (L) and the milliliter (mL).
  - A liter is a cube 1 dm (10 cm) long on each side.
  - A milliliter is a cube 1 cm long on each side.



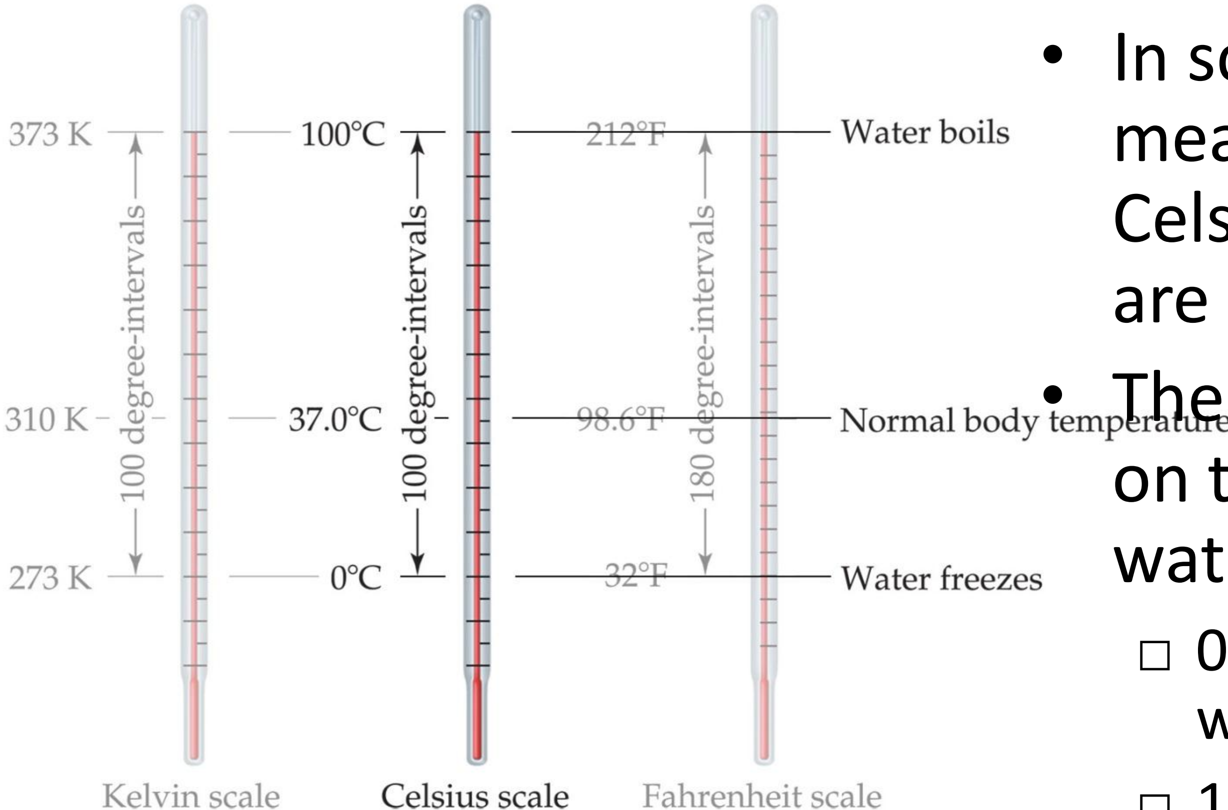
Temperature:



proportional to the average kinetic energy of the particles in a sample.

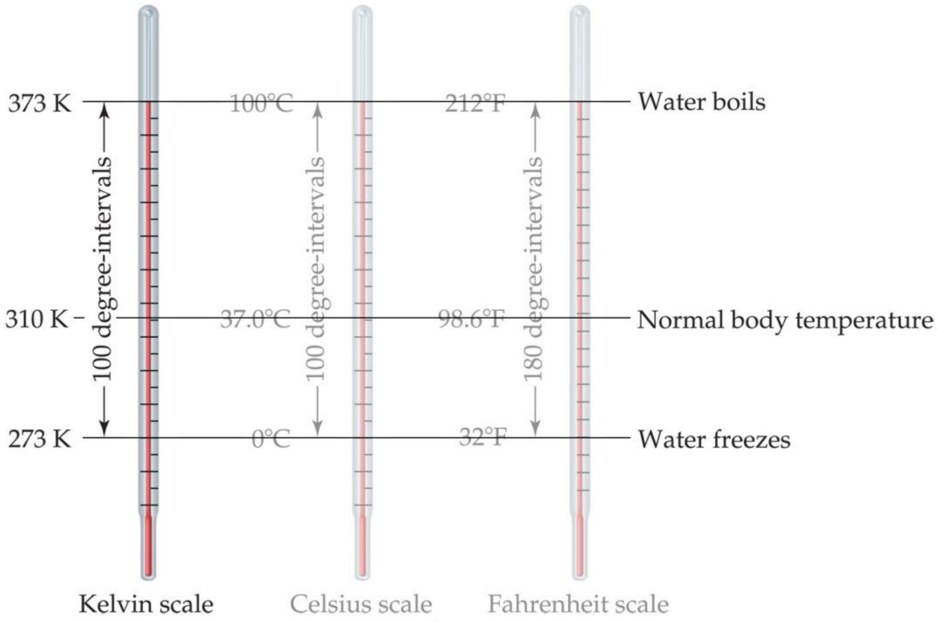
K.E. = $1/2mv^2$

Temperature



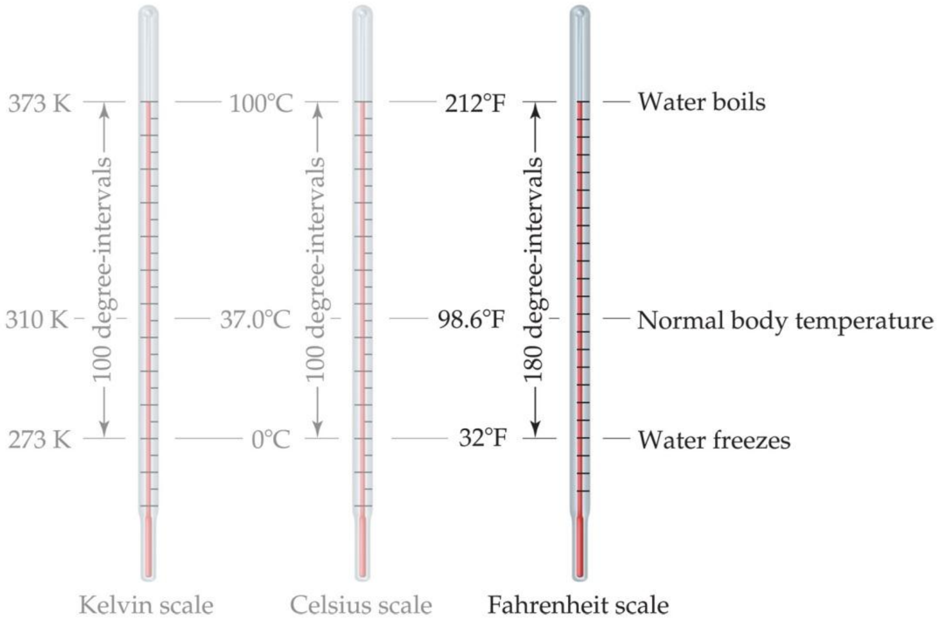
- In scientific measurements, the Celsius and Kelvin scales are most often used.
- The Celsius scale is based on the properties of water.
  - 0°C is the freezing point of water.
  - 100°C is the boiling point of water.

Temperature



- The Kelvin is the SI unit of temperature.
- It is based on the properties of gases.
- 0 K = 0 K.E.
- There are no negative Kelvin temperatures.
- K = °C + 273.15

Temperature



- The Fahrenheit scale is not used in scientific measurements.
- °F = 9/5(°C) + 32
- °C = 5/9(°F) − 32

Density:

Physical property of a substance

Intensive.

$$d = \frac{m}{V}$$

Density of selected substances

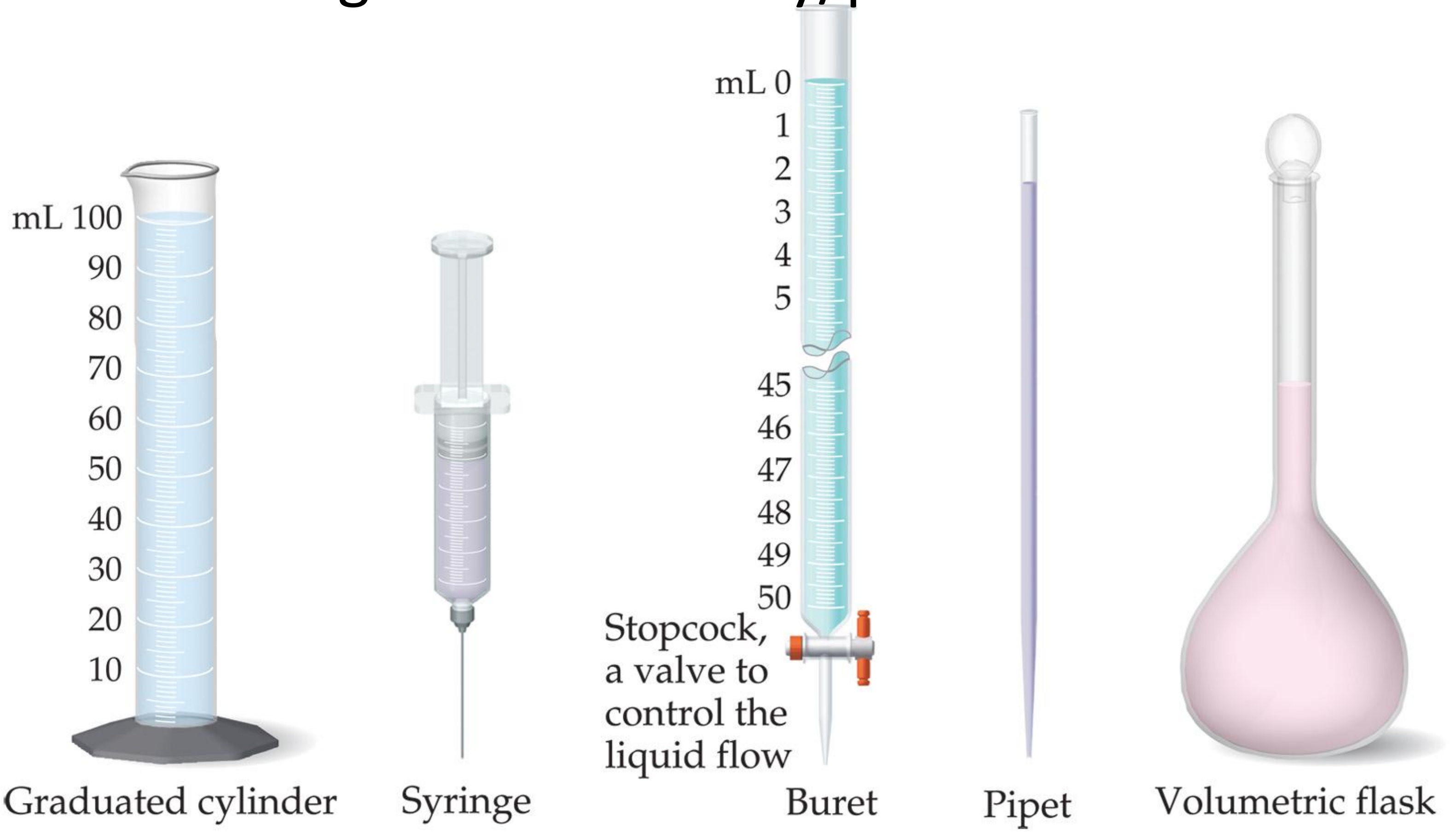
TABLE 1.6 Densities of Some Selected Substances at 25°C

Substance	Density (g/cm^3)
Air	0.001
Balsa wood	0.16
Ethanol	0.79
Water	1.00
Ethylene glycol	1.09
Table sugar	1.59
Table salt	2.16
Iron	7.9
Gold	19.32

Uncertainty in Measurement

Uncertainty in Measurements

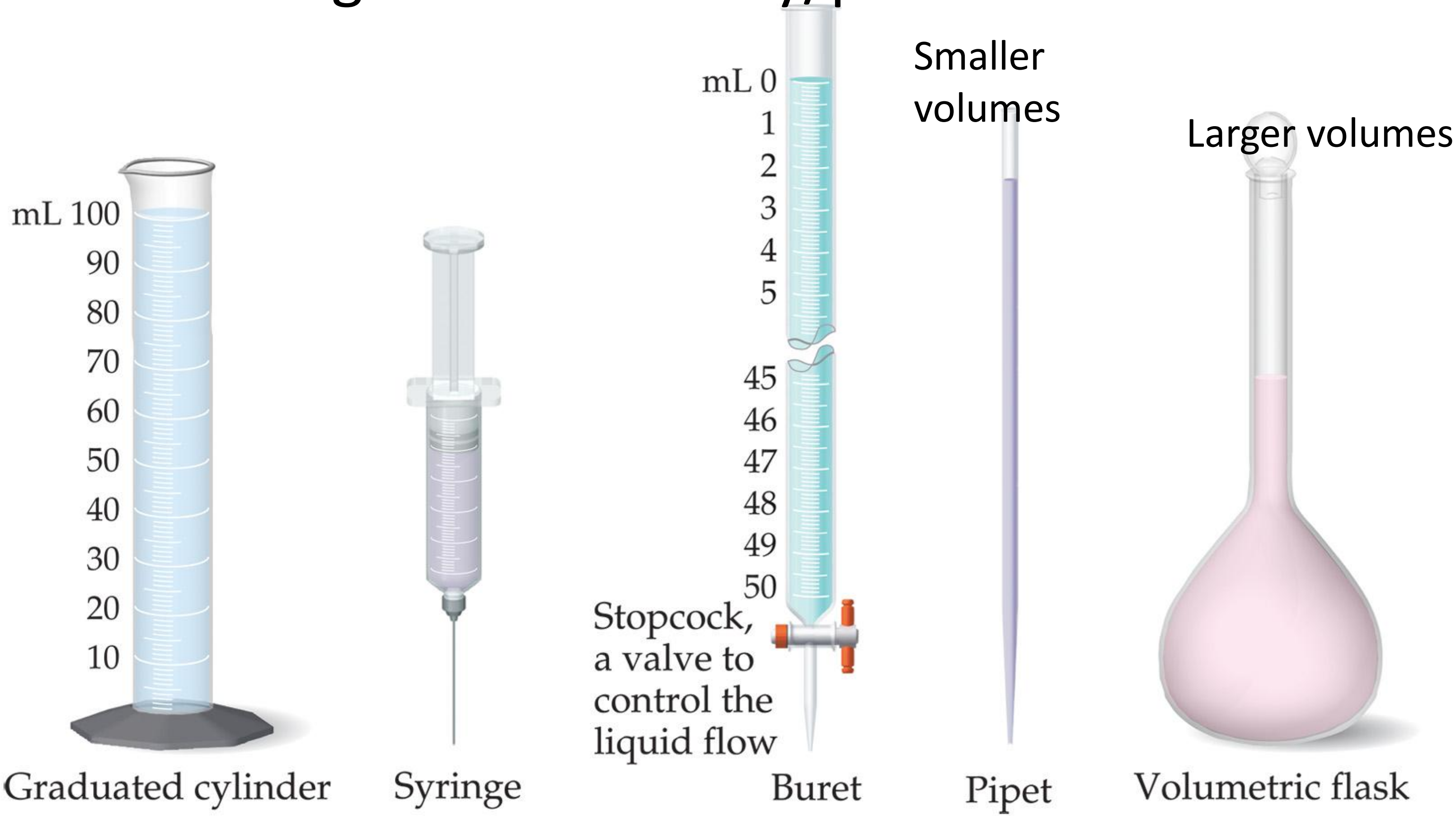
Different measuring devices have different uses and different degrees of accuracy/precision.



Which are more accurate?

Uncertainty in Measurements

Different measuring devices have different uses and different degrees of accuracy/precision.



It depends on amount

Exact versus inexact numbers

Exact	**Inexact**
1000 g/kg	ruler measure
2.54 cm/in	Temp. reading
12/dozen	volume or mass
any conversion factor	etc.

Significant Figures

- The term significant figures refers to digits *that were measured*.
- When rounding calculated numbers, we pay attention to significant figures so we do not overstate the precision of our answers.

Significant Figures

1. All nonzero digits are significant. (sig figs in **red**)

   423.444

2. Zeroes between two significant figures are themselves significant.

   42,300045   42,340.0025

3. Zeroes at the beginning of a number are never significant.

   00042345.0 0.00048

4. Zeroes at the end of a number are significant if a decimal point is written in the number.

   423,000 versus: 423,000. or: 423,000.000

Significant Figures

- When addition or subtraction is performed, answers are rounded to the least significant decimal place.

```
  24.245
+22.33488
  46.57988 = 46.580
```

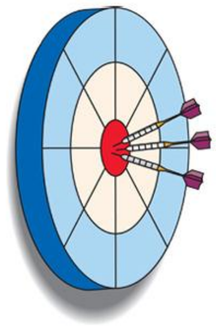
- When multiplication or division is performed, answers are rounded to the number of digits that corresponds to the *least* number of significant figures in any of the numbers used in the calculation.

```
  35.8750      (6 sig figs)
 X40.006800  (8 sig figs
1435.24395 = 1435.24  (6 sig figs)
```

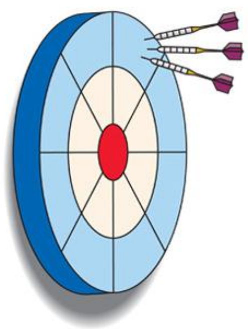
Accuracy versus Precision

- Accuracy How close a measurement is to the true value. (How right you are)

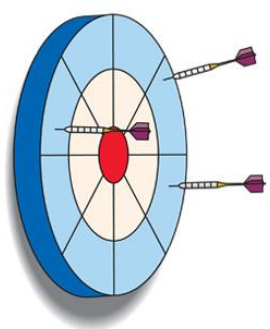
- Precision How close measurements are to each other. (Reproducibility). Precise but incorrect data are often the result of systematic errors.



Good accuracy
Good precision



Poor accuracy
Good precision



Poor accuracy
Poor precision

Dimensional analysis

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What do virtually all problems in chemistry have in common?

Dimensional analysis

Convert centimeters to feet: 1 cm = ? feet

Know: 2.54 cm = 1 in, 12 in = 1 foot.

$$\frac{1in}{2.54cm}\left(\frac{1ft}{12in}\right) = 0.032\frac{ft}{cm}$$

Dimensional Analysis

- What do I need on top?
- What do I need on the bottom?
- What do I know?
- How do I get there?
- Note: You will always be given the conversion factors you need, you don't have to memorize them.

Dimensional analysis, examples

The speed of light is $2.998x10^{10}$ cm/s. What is it in km/hr?

Know: 1 km =1000m, 1m=100cm 60 min =1 hr, 60 sec =1 min

What do I need on top? *kilometers*

What do I need on the bottom? *hours*

$$2.998x10^{10}\frac{cm}{s}\left(\frac{1m}{100cm}\right)\left(\frac{1km}{1000m}\right)\left(\frac{60\sec}{1\min}\right)\left(\frac{60\min}{1hr}\right)=1.0892x10^{9}\,km/hr$$

Dimensional analysis, examples

The Vehicle Assembly Building (VAB) at the Kennedy Space Center has a volume of: 3,666,500m^3. What is it in liters?

Know: 1 L = 1 dm^3, 1dm = 0.1 m

What do I need on top? *Liters*

What do I need on the bottom? *nothing*

$$3,666,500m^3 = \left(\frac{dm}{0.1m}\right)^3 \left(\frac{1L}{1dm^3}\right) = 3.6665x10^9 L$$

Dimensional analysis, examples

An individual suffering from high cholesterol has 232 mg cholesterol per 100.0 mL of blood. How many grams of cholesterol in the blood, assuming a blood volume of 5.2 L?

Know: 1 L = 1000 mL, 1g = 1000mg

What do I need on top? *grams*

What do I need on the bottom? *patient*

$$232\frac{mg}{100.0mL}\left(\frac{1000mL}{1L}\right)\left(\frac{5.2Lblood}{patient}\right)\left(\frac{1g}{1000mg}\right)=12.\frac{g}{patient}$$