

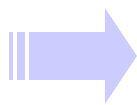
Kotz and Treichel Introduction and Chapter 1

Basic Concepts of Chemistry

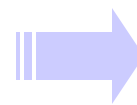
The Tools of Quantitative
Chemistry

Introduction - What is Chemistry

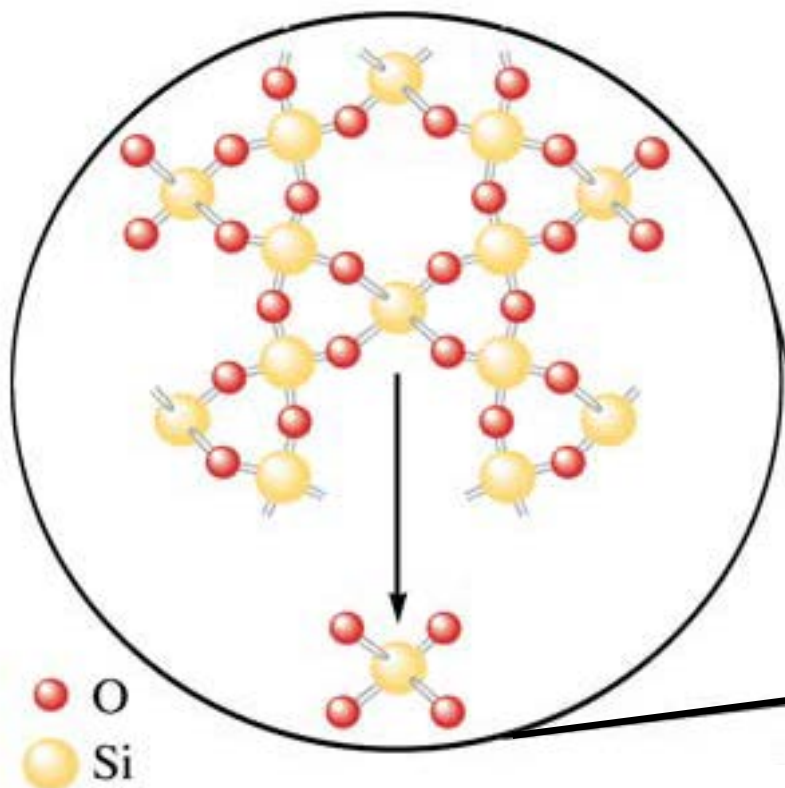
atomic /
molecular
level



microscopic
level



macroscopic
level



The Scientific Method

Make observations

Collect Data

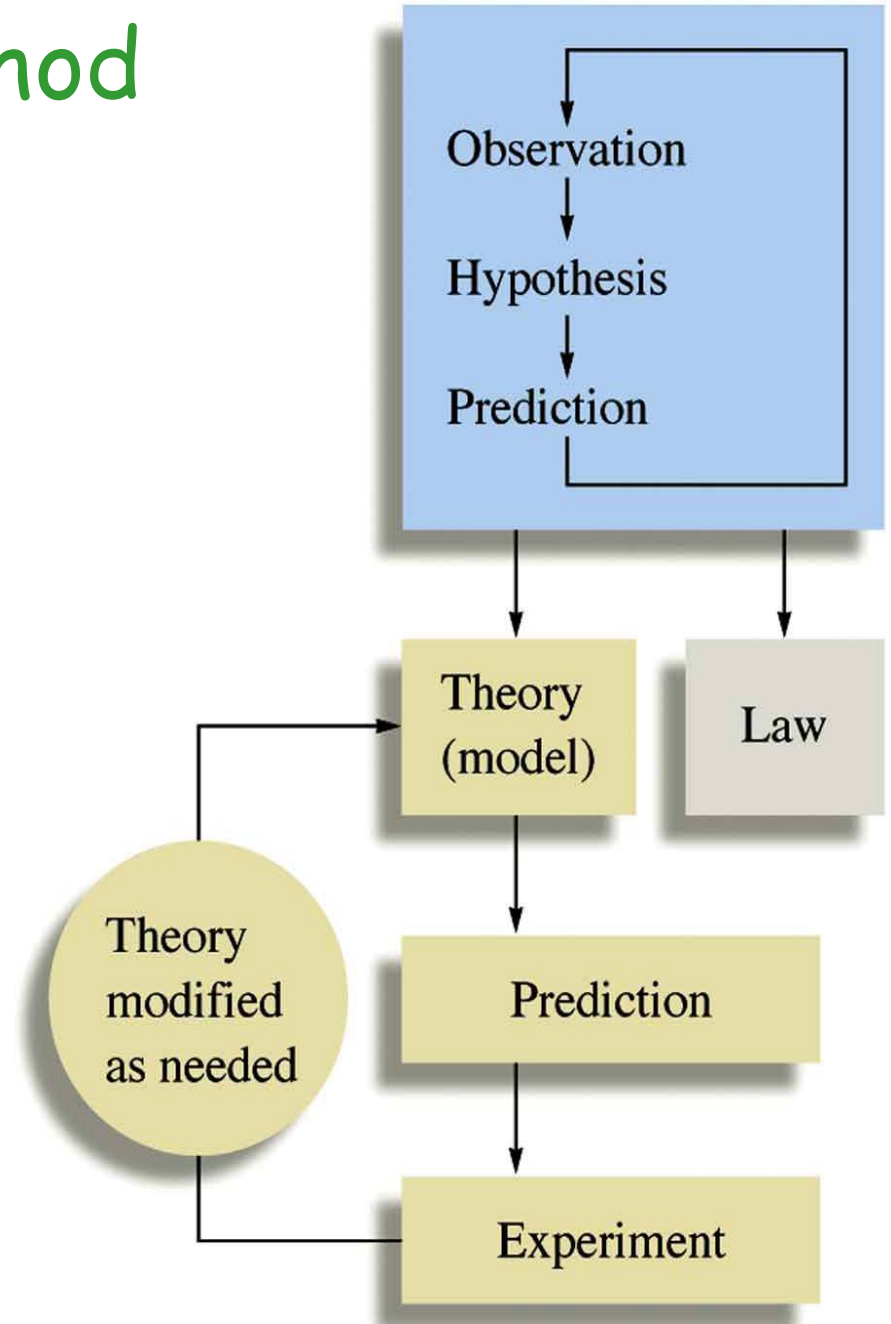
Formulate a hypothesis

A tentative explanation

Test the hypothesis

Make a prediction

Do an experiment to test the prediction.



The Scientific Method

Observation - Example

A gas, for example, air can be compressed: the higher the pressure, the smaller the volume.

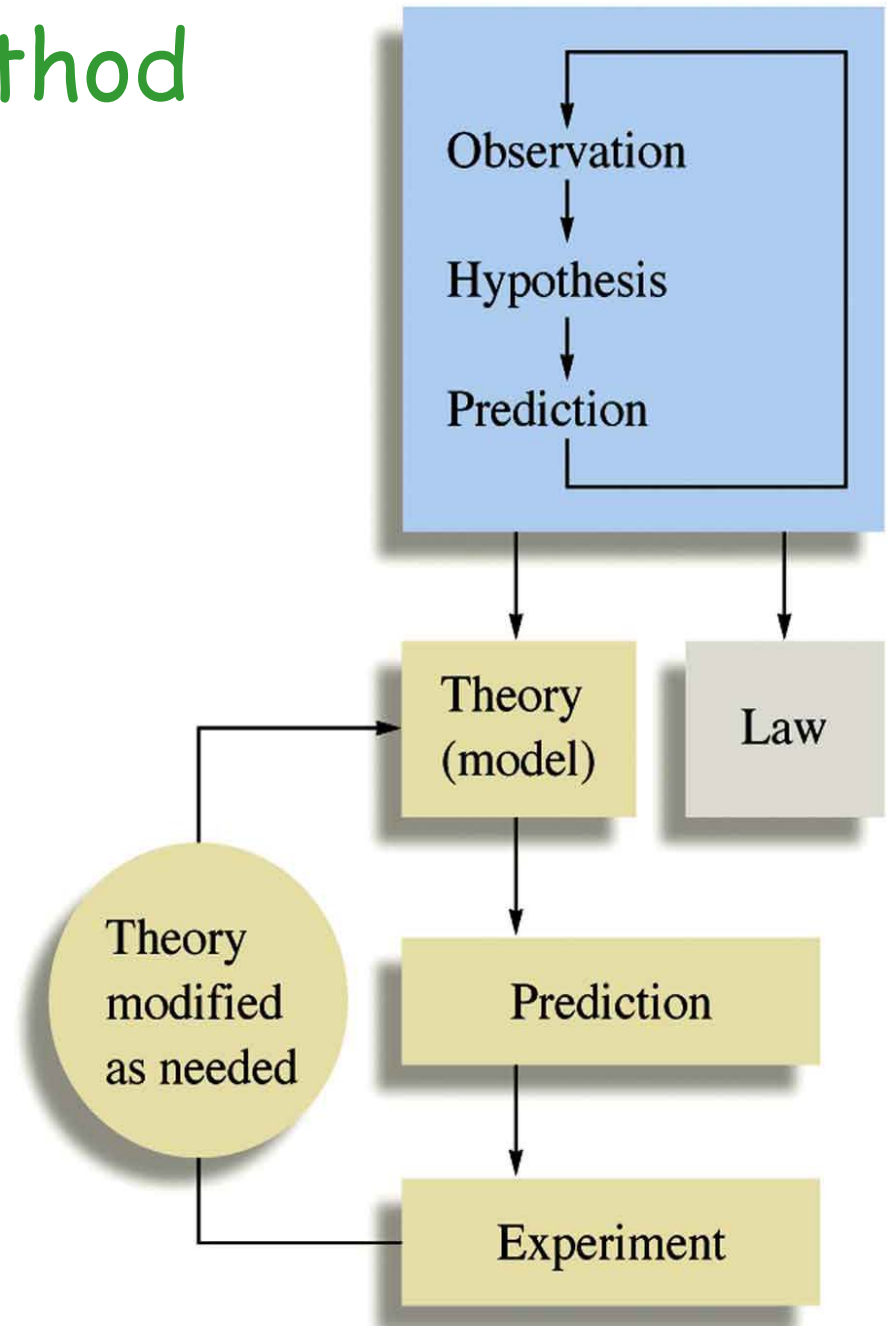
Hypothesis:

Gases are made of molecules with a lot of space between.

Prediction and Experiment:

All gases will behave the same way.

Try compression experiments with hydrogen, methane, helium, etc.



The Scientific Method

Law

The product, pressure times volume, is a constant, if the temperature is not changed.

Theory (Model)

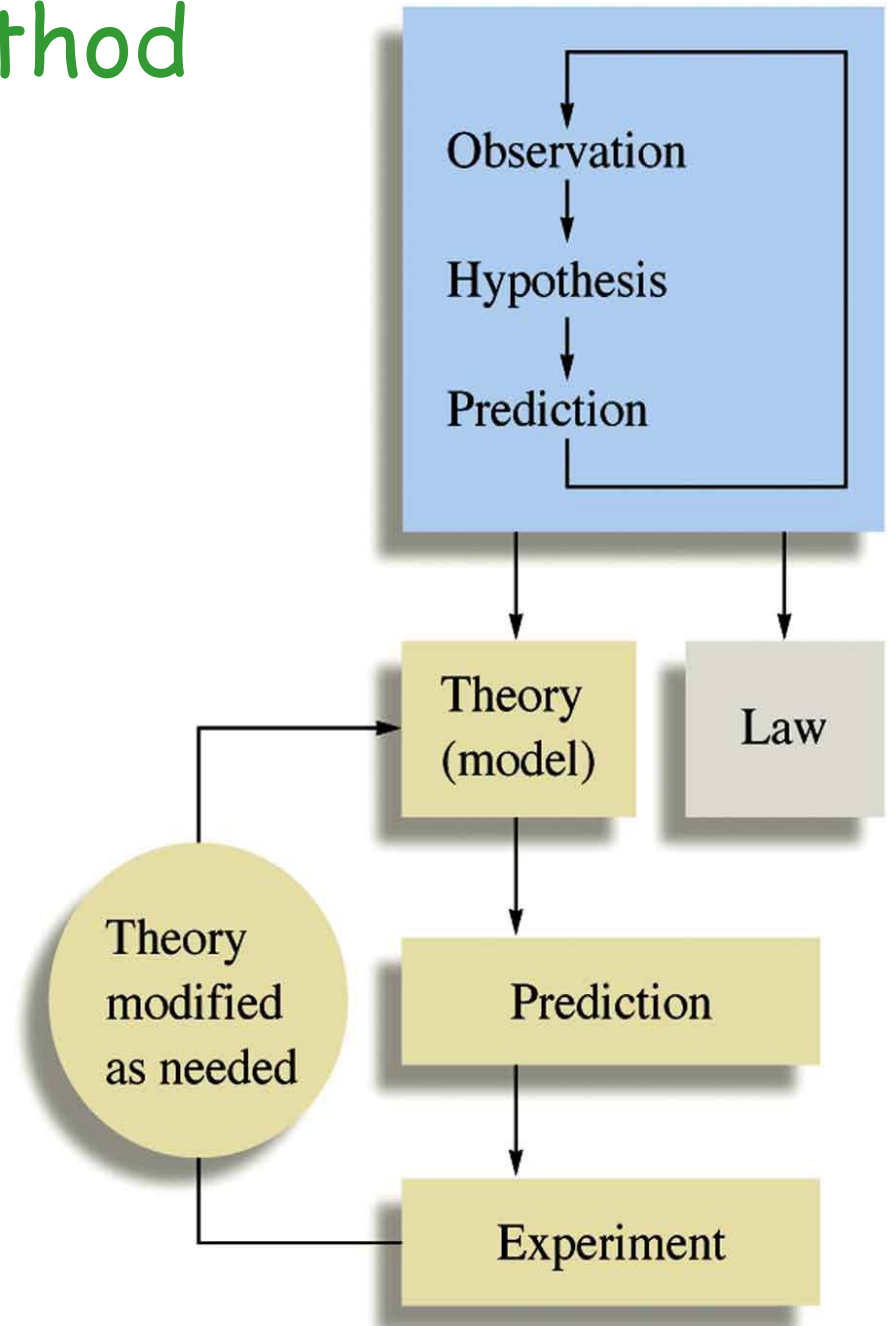
The simple "kinetic theory" of ideal gases gives the hypothesis and the law a quantitative basis.

Prediction

of the physical properties of real gas samples.

Improve the model

to account for real gas behaviour.



The Scientific Method

Serendipity

There are often lucky breaks. The discovery of penicillin and teflon, for example.

Integrity in Science

It is all too easy to become so fond of a pet theory to the point where we design experiments to "test" the theory which are biased to support it.

We also have to be concerned about the uses to which our discoveries may be put. We need not only to make the discoveries, but also discover the controls on them.

Sustainability and "green chemistry"

- Not harm the environment;
- Not use the natural resources faster than nature can regenerate them

States of Matter

Kinetic-Molecular Theory

- **States of matter:**

- **Solids**: hard, fixed shape and volume, little volume change with changes of temperature;
- **Liquids**: fluid, no fixed shape (takes the shape of the container), fixed volume, little little volume change with changes of temperature(a little more than solids however);
- **gases**: fluid, no fixed shape or volume (expands in all available space), volume and pressure change dramatically with temperature changes;
- **condensed matter**: solids and liquids since they occupy a minimum volume

At very low temperature, everything is solid (there is one exception, He).

As temperature increases: solid → liquid → gas

States of Matter

Kinetic-Molecular Theory

- **Kinetic-molecular theory**

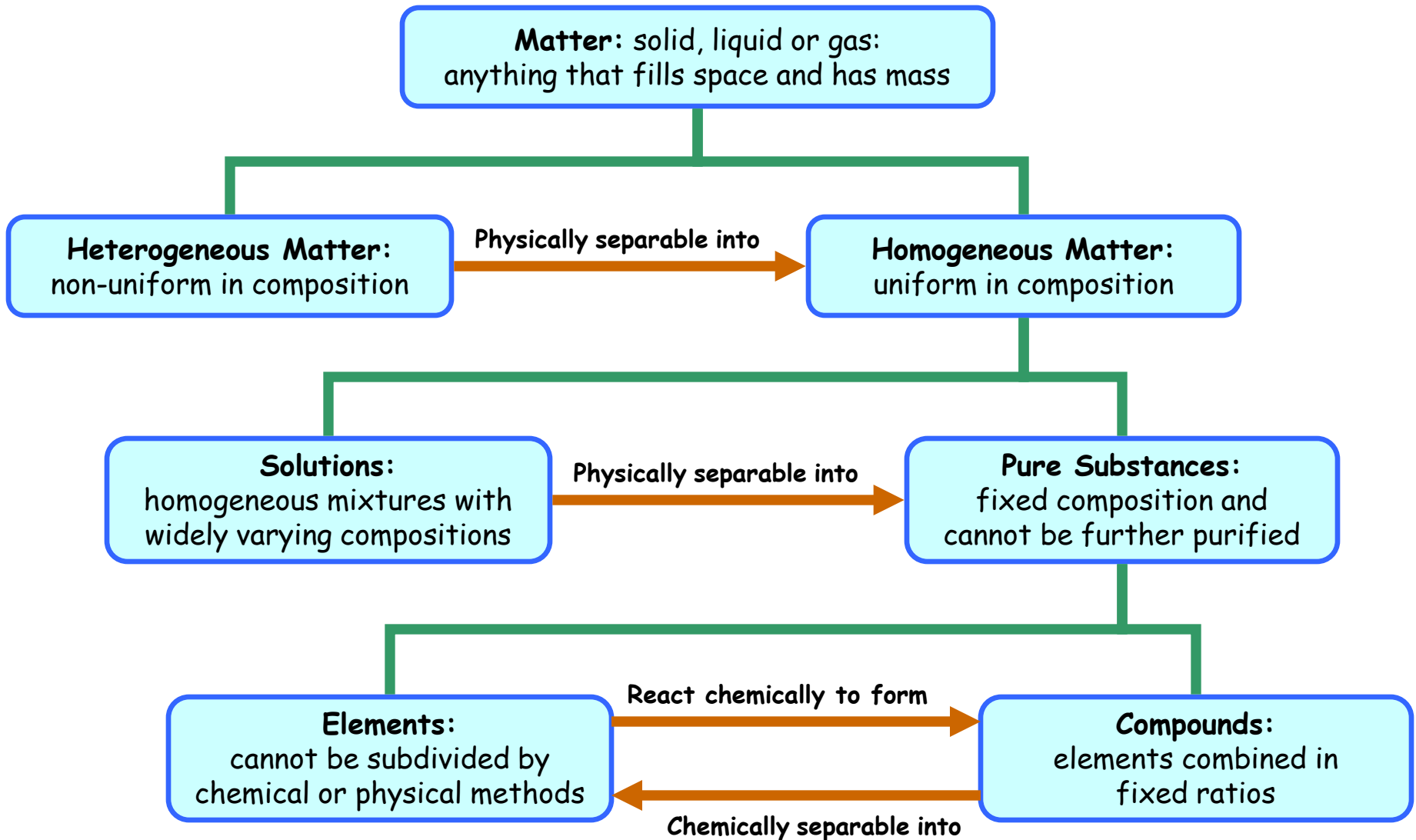
- * All particles are in motion;
- * The higher the temperature, the faster the motion;
- * The energy of the motion is called “kinetic energy”;
- * The kinetic energy is responsible for the temperature.

- **Solids**: particles (atoms/ions/molecules) are packed closely together in a regular array; they vibrate around their average position; very low compressibility;

- **Liquids**: random orientation of the particles relative to one another; they occupy a little more space and they keep moving relative to one another; they still stay close to one another; very low compressibility;

- **gases**: particles are far apart; particles move fast in a linear motion until they collide with another particle, when kinetic energy is exchanged and direction of motion changes; very high compressibility.

Classifying Matter



Matter

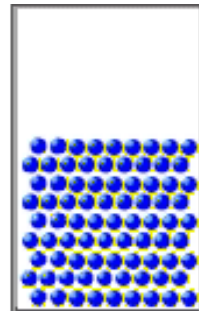
Ice

Lead

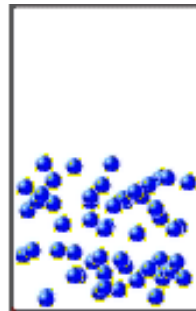


Solids have a definite shape and more or less fixed volume. They do not flow except under the most extreme conditions of shear or stretching forces.

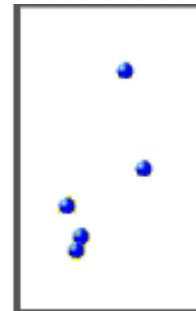
At the molecular level, the atoms ions or molecules vibrate in fixed positions and orientations with respect to each other.



Solid



Liquid

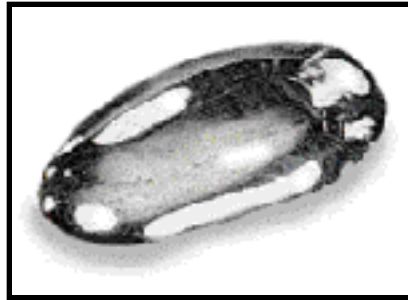


Gas

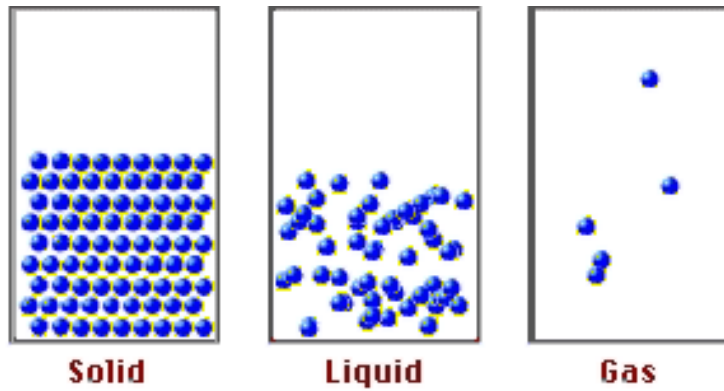
Matter

Liquids are fluid and can take any shape, but unlike a gas, they have a more or less fixed volume. In any container that has a greater volume than the liquid, there will be a surface.

Mercury



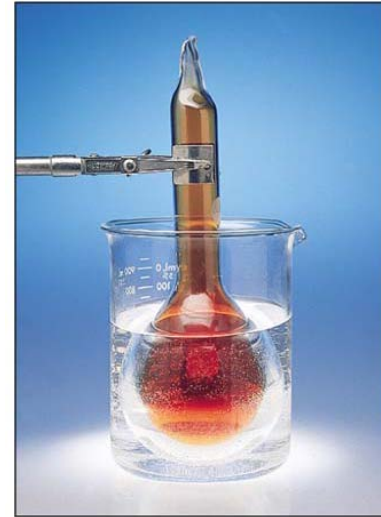
At the molecular level, the atoms, ions or molecules are in contact with each other, but not in fixed positions and orientations with respect to each other: they are in continual motion.



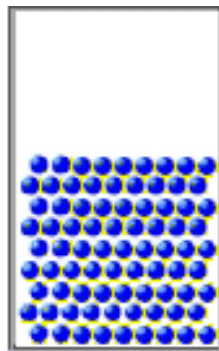
Matter

Gases are fluid with no definite shape or volume. They fill any container completely and homogeneously by expanding freely.

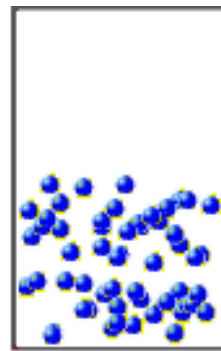
At the molecular level, the atoms or molecules of the gas are in continual random motion and their average separation depends only on the number of molecules and the volume of the container.



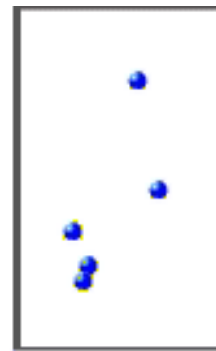
Nitrogen dioxide, NO_2



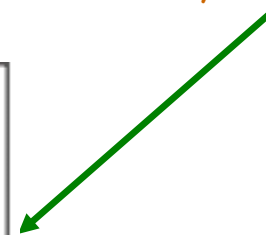
Solid



Liquid



Gas



Mixtures



Homogeneous Mixture
(e.g. a solution)

Heterogeneous Mixture
(e.g. mixture of particles)

Heterogeneous mixtures are not always easy to recognize. There are many things which are only seen to be heterogeneous under a microscope, e.g.:

Liquids such as: milk, blood

Solids such as: brass, steel

Mixtures and Pure Substances

It is rare to obtain pure substances in the laboratory or in an industrial setting. One of the main activities of chemists working on syntheses, for example, is the separation of the desired product from unwanted by-products.

Several methods can be used to separate mixtures into their component pure substances, for example:

Filtration

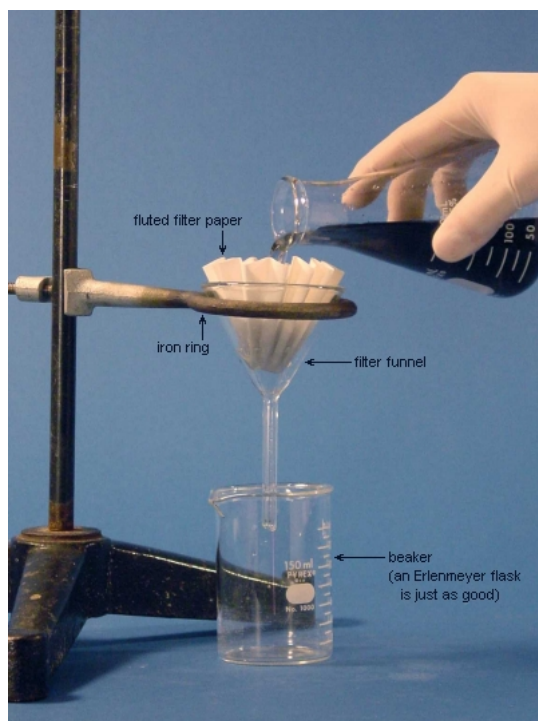
Distillation

Chromatography

But, any method may fail, and it is not always possible to be sure that a pure substance has, in fact, been isolated.

For example, distillation of a mixture of ethanol and water cannot remove the last 5% or so of the less volatile water from the alcohol.

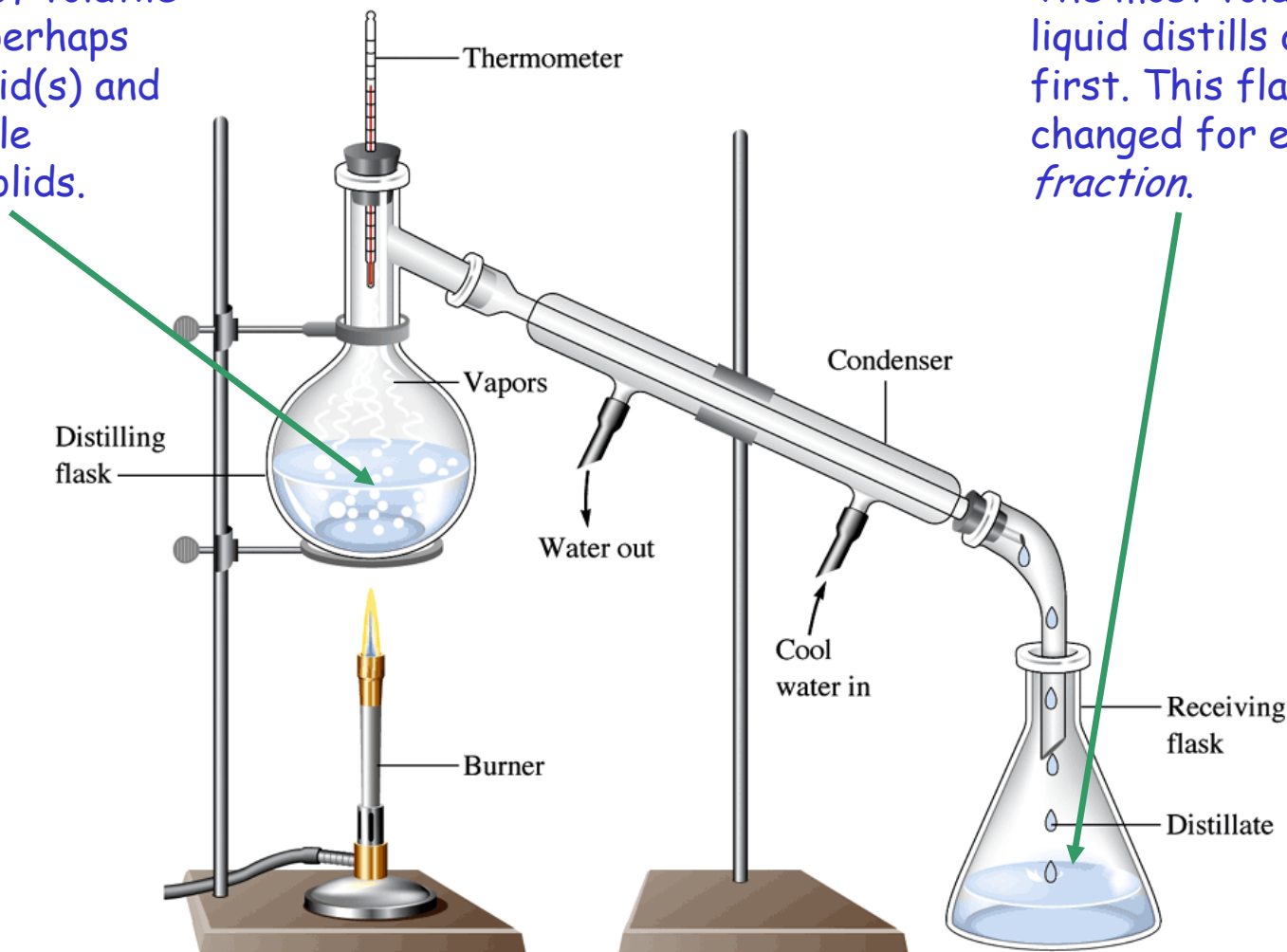
Separation of a Mixture by Filtration



<http://www.chem.ucalgary.ca/courses/350/laboratory/expt03gf.html>

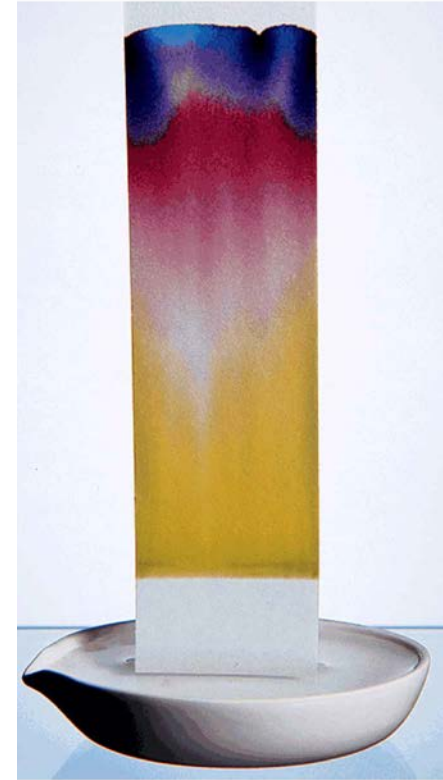
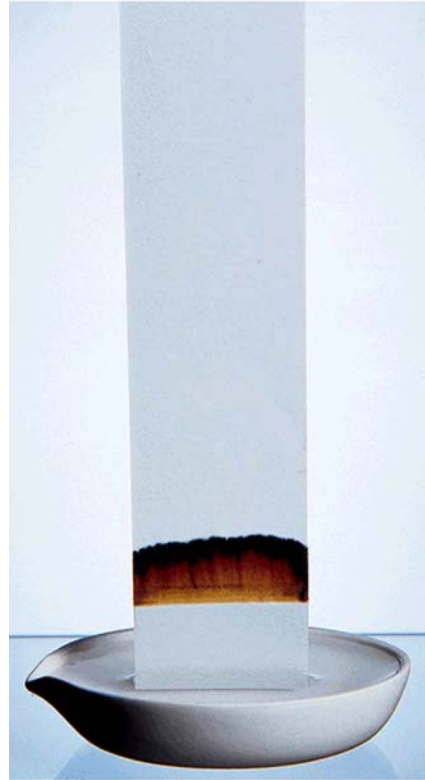
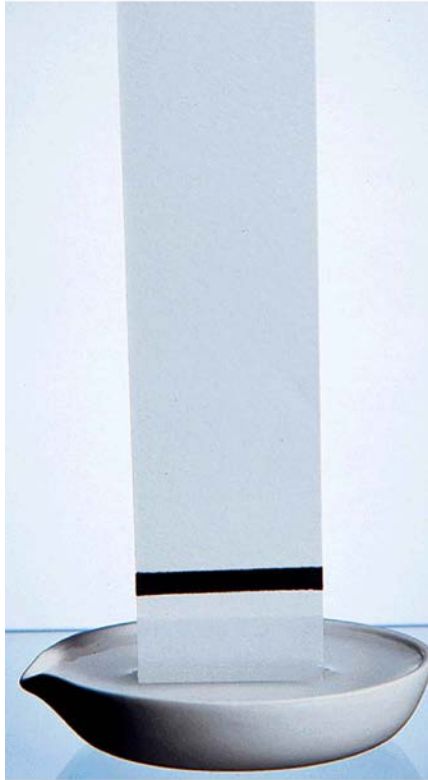
Separation of a Mixture by Distillation

A mixture of volatile liquids, or perhaps volatile liquid(s) and involatile dissolved solids.



The most volatile liquid distills over first. This flask is changed for each fraction.

Separation of a Mixture by Chromatography



Pictures taken from Zumdahl

Analyte: a mixture of coloured dyes (black ink)

Stationary Phase: filter paper (porous paper)

Moving phase: probably a mixture of water and alcohol

Elements and Atoms - Compounds and Molecules

One way to separate a compound into its elements - electrolysis

The *compound* is water, a *molecule* made from two *atoms* of hydrogen and one atom of oxygen.

Water can be separated into its constituent elements by *electrolysis*, giving molecules consisting of only one kind of atom, hydrogen, H_2 and oxygen, O_2

The reverse of this process, the production of water from hydrogen and oxygen in a *fuel cell*, can generate electricity.

Oxygen, O_2

Hydrogen, H_2

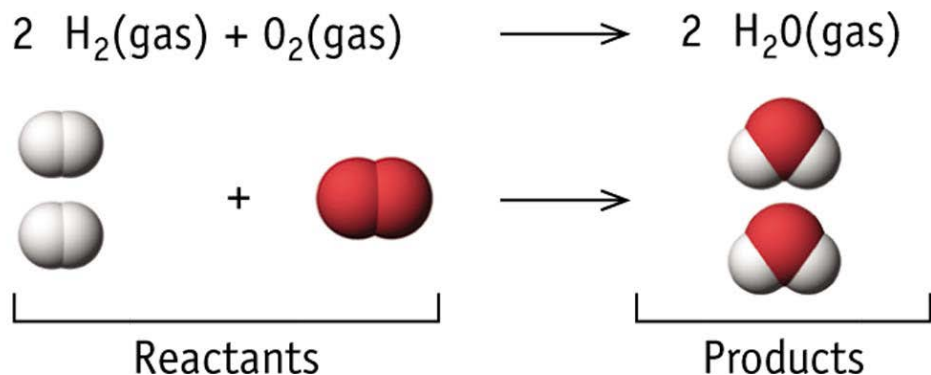


Water, H_2O

Elements and Atoms - Compounds and Molecules

The elements re-combine to form a
chemical compound

We show how substances react together by
writing a *chemical equation* to describe the
process:



As well as a summary of what happens in the reaction, we can use these
equations to do *stoichiometric calculations* - work out what quantities if
each *reactant* are required and how much *product* is produced.

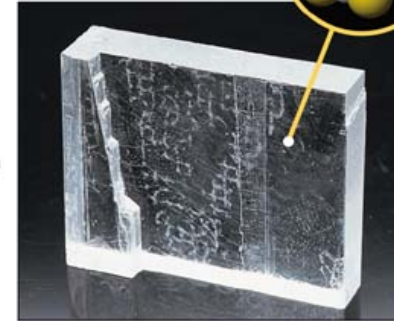
Elements and Atoms - Compounds and Molecules

Sodium
Na (a metal)



+

Chlorine
Cl₂ (a gas)



Sodium Chloride, NaCl
(an ionic solid)



Note that neither sodium nor sodium chloride have identifiable molecules

Physical Properties

<u>Property</u>	<u>Using the property to Distinguish Substances</u>
Colour	Is the substance coloured or colourless? What is the colour, and what is its intensity?
State of matter	Is it a solid, liquid or gas? If it is a solid, what is the shape of its particles?
Melting point	At what temperature does a solid melt?
Boiling point	At what temperature does a liquid boil?
Density	What is its density (mass per unit volume)?
Solubility	What mass of the substance can dissolve in a given volume of water or some other solvent?
Electric conductivity	Is it a conductor of electricity or an insulator?
Malleability	How easily can a solid be deformed?
Ductility	How easily can a solid be drawn into a wire?
Viscosity	How easily does a liquid flow?

Physical Properties - Density

Density, mass per unit volume, is a type of *conversion factor*.

$$\text{Density} = \frac{\text{Mass}}{\text{Volume}}$$

We most commonly use density in the laboratory to convert a quantity called for by a particular procedure from something which may be inconvenient to measure out, to something easier.

For example, you might be following a procedure which calls for 200. g of toluene, a liquid, in some reaction to make a derivative of it. For liquids, it is usually easier to measure by volume:

The density of toluene is 0.867 g cm⁻³ or g mL⁻¹ or g/cm³ or g/mL

$$\frac{200. \text{ (g)}}{\text{required volume (mL)}} = 0.867 \text{ g/mL}$$

$$\text{So the required volume} = \frac{200. \text{ (g)}}{0.867 \text{ (g/mL)}} = 231. \text{ mL}$$

Extensive and Intensive Properties

Extensive properties: depend on the amount of substance.

Examples: mass, volume, amount of heat obtained by combustion

Intensive properties: do not depend on the amount of substance.

Examples: melting point, boiling point, density

Pay attention: The ratio of two extensive properties is an intensive property

Example: $\text{mass/volume} = \text{density}$

Physical and Chemical Changes

Physical Changes: Changes in physical properties, no change in the identity of a substance

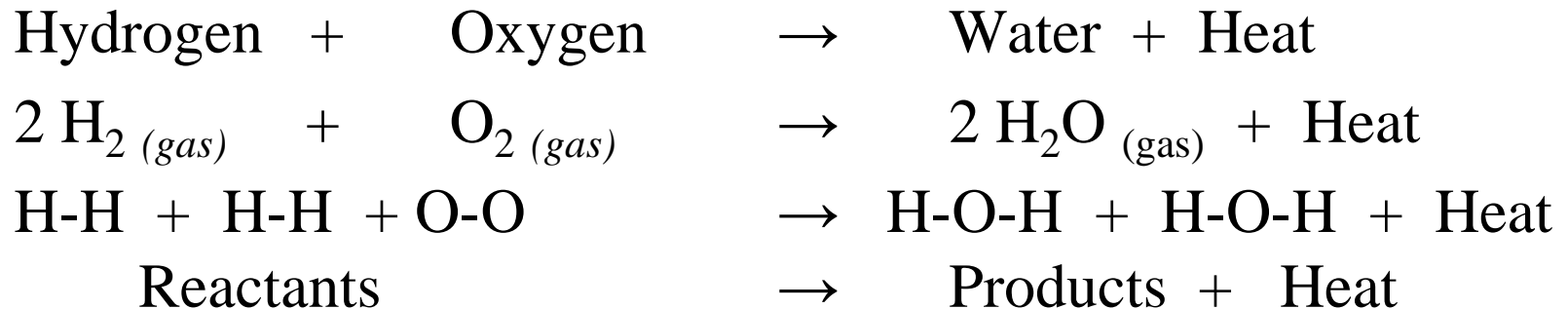
From	To	Physical change	Example
Solid	Liquid	Melting	Ice melting
Solid	Gas	Sublimation	Moth balls smell
Liquid	Gas	Vaporization	Water evaporating
Gas	Liquid	Liquefaction	Fog, mist on a cold surface
Liquid	Solid	Solidification	Water freezing
Gas	Liquid or solid	Condensation	

(Liquids and solids are called “condensed matter”)

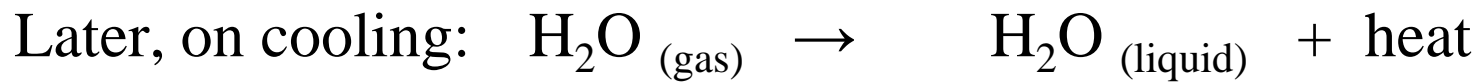
Physical and Chemical Changes

Chemical changes: There is a change in the identity of substances

1. Combustion: Chemical reaction:



2. Liquefaction of steam to water: Physical change



Energy: Some basic principles

Energy can be classified in two kinds:

- **Kinetic energy:** associated with motion
 - - **Thermal energy:** motion of submicroscopic particles (atoms, ions, molecules): mostly vibrations and rotations
 - - **mechanical energy:** motion of macroscopic objects (balls, cars, people)
 - - **electrical energy:** motion of electrons in conductors
 - - **acoustic energy:** pressure waves in materials (transmission of sound)
- **Potential energy:** associated with position or status of objects
 - - **gravitational energy:** attraction by very large bodies (Earth, Sun)
 - - **Electrostatic energy:** energy associated with the separation of electrical charges
 - - **chemical energy:** stored in chemical bond. Chemical reactions involve an absorption or emission of energy due to bonds being broken and others being formed
 - - **nuclear energy:** energy/mass conversion (Einstein relativity theory)

Conservation of Energy

- **Law of conservation of energy:** Energy can neither be created nor destroyed i.e. the total energy of the universe is constant. (including the energy changes occurring at nuclear reactions)

In a chemical reaction:

Sum of the bond energy in all reactants + heat absorbed = Sum of the bond energy in all products, for **endothermic reactions** (heat is absorbed)

or

Sum of the bond energy in all reactants = Sum of the bond energy in all products + heat emitted, for **exothermic reactions** (heat is emitted)

Note: For nuclear reactions, the energy change due to change of mass must also be taken into account: the total energy (including in the mass) of the universe is constant

Let's Review

The tools of Quantitative Chemistry

Qualitative and quantitative information

- **Qualitative information**

- nonnumerical observations
- **Examples:** Color of a substance, its physical appearance (hard solid, soft material, grainy, thick liquid like oil, non-thick liquid like water, gas, smell, feel of temperature: cold, warm, hot)

- **Quantitative information**

- numerical observations: has a number and a unit, such as: the length of an object is 5.3 cm
- **Examples:** length, mass, volume, temperature

Units of Measurement

Table 1.2 • Some SI Base Units

Measured Property	Name of Unit	Abbreviation
Mass	kilogram	kg
Length	meter	m
Time	second	s
Temperature	kelvin	K
Amount of substance	mole	mol
Electric current	ampere	A

Units of Measurement

The candela, symbol cd, is the unit of
luminous intensity

Units of Measurement

There are a number of common derived units:

Volume	litre	L	1000 cm^3
Force	newton	N	1 kg m sec^{-2}
Pressure	pascal	Pa	1 N m^{-2}
Energy	joule	J	$1 \text{ kg m}^2 \text{ s}^{-2}$

Table 1.3 • Selected Prefixes Used in the Metric System

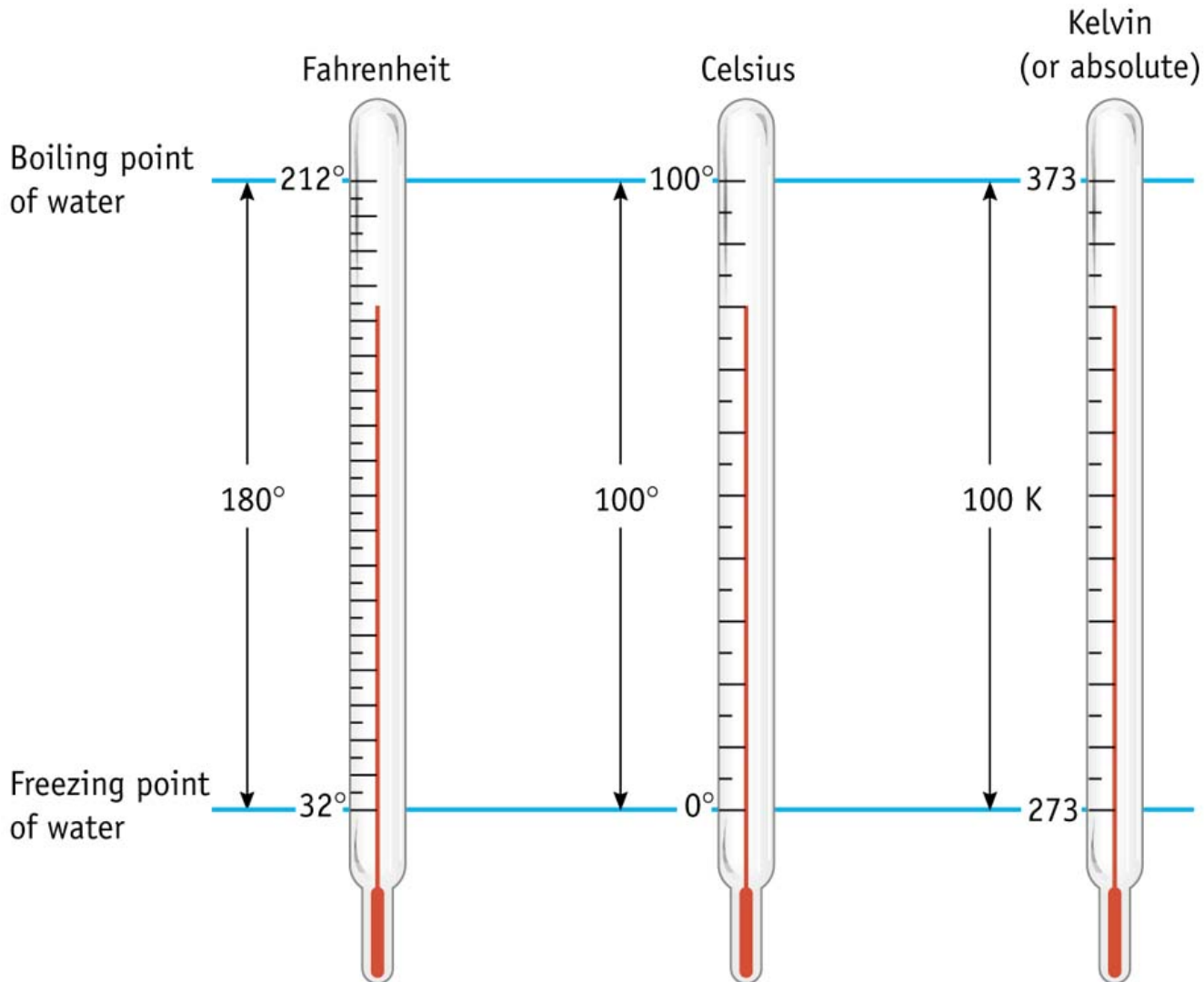
Prefix	Abbreviation	Meaning	Example
mega-	M	10^6	1 megaton = 1×10^6 tons
kilo-	k	10^3	1 kilometer (km) = 1×10^3 m
deci-	d	10^{-1}	1 decimeter (dm) = 1×10^{-1} m
centi-	c	10^{-2}	1 centimeter (cm) = 1×10^{-2} m
milli-	m	10^{-3}	1 millimeter (mm) = 1×10^{-3} m
micro-	μ	10^{-6}	1 micrometer (μm) = 1×10^{-6} m
nano-	n	10^{-9}	1 nanometer (nm) = 1×10^{-9} m
pico-	p	10^{-12}	1 picometer (pm) = 1×10^{-12} m

Selected prefixes used in the metric system (continued)

Femto f 10^{-15}

Example: 1 femtometer (fm) = 1×10^{-15} m

Physical Properties - Temperature



Physical Properties - Temperature Scales

Fahrenheit

Zero on this scale was set to the lowest temperature attainable with an ice/salt mixture, and 100 °F was Fahrenheit's usual body temperature. Nowadays the fixed-points are the freezing point of pure water, 32 °F and the boiling point of water, 212 °F.

Celsius (formerly centigrade)

The fixed points are the freezing point of pure water, 0 °C, and the boiling point of water, 100 °C. This is the "everyday" temperature scale used in the lab.

Kelvin (or absolute)

The size of the degree is the same as Celsius, but 0 K is set to the lowest attainable temperature, currently believed to be -273.15 °C. This temperatures expressed in Kelvins *must* be used in many calculations, for example those involving the effect of a change in temperature on the volume or pressure of a gas sample. **To convert °C to K, just add 273.15.**

Physical Properties - Temperature Scales

Converting degrees Fahrenheit to degrees Celsius and vice versa

T_F = temperature in degrees Fahrenheit

T_C = temperature in degrees Celsius

T_K = temperature on the Kelvin scale

$$T_C = (T_K - 273.15 \text{ K}) \times 1^\circ \text{C}/1 \text{ K}$$

$$T_K = (T_C \times 1 \text{ K}/1^\circ \text{C}) + 273.15 \text{ K}$$

$$T_C = (T_F - 32^\circ \text{F}) \times 5^\circ \text{C}/9^\circ \text{F}$$

$$T_F = (T_C \times 9^\circ \text{F}/5^\circ \text{C}) + 32^\circ \text{F}$$

Problem 1: Liquid nitrogen boils at 77 K. What is this temperature on the Fahrenheit scale?

$$T_C = (T_K - 273.15 \text{ K}) \times 1^\circ \text{C}/1 \text{ K} = 77 - 273.15 = -196.15^\circ \text{C}$$

$$T_F = (T_C \times 9^\circ \text{F}/5^\circ \text{C}) + 32^\circ \text{F} = (-196.15^\circ \text{C} \times 9^\circ \text{F}/5^\circ \text{C}) + 32^\circ \text{F} = -321.07^\circ \text{F} = -321^\circ \text{F}$$

Problem 2: A person has a body temperature of 102.5°F. What is this temperature on the Celsius scale and on the Kelvin scale?

$$T_C = (T_F - 32^\circ \text{F}) \times 5^\circ \text{C}/9^\circ \text{F} = (102.5^\circ \text{F} - 32^\circ \text{F}) \times 5^\circ \text{C}/9^\circ \text{F} = 39.2^\circ \text{C}$$

$$T_K = T_C + 273.15 \text{ K} = (39.2^\circ \text{C} \times 1 \text{ K}/1^\circ \text{C}) + 273.15 \text{ K} = 312.32 \text{ K} = 312.3 \text{ K}$$

Length, Volume and Mass

- Distances at the molecular scale: nanometers or picometers

Exemple: The O-H distance in the water molecule is
 $95.8 \text{ pm} = 0.0958 \text{ nm} = 9.58 \times 10^{-11} \text{ m}$

$1 \text{ \AA} \text{ (Angström)} = 1 \times 10^{-10} \text{ m}$ (not a SI unit, used widely in the past for atomic sizes since 1 \AA is about the size of an atom).

$1 \text{ \AA} = 0.1 \text{ nm} = 100 \text{ pm}$

Length, Volume and Mass (continued)

Length: $1 \text{ m} = 100 \text{ cm}$

Surface area: $1 \text{ m}^2 = (100 \text{ cm})^2 = 10\,000 \text{ cm}^2$

Volume: $1 \text{ m}^3 = (100 \text{ cm})^3 = 1\,000\,000 \text{ cm}^3$

1 L (liter) = 1 dm³ by definition of the liter.

$10 \text{ cm} = 1 \text{ dm} \Rightarrow 1 \text{ dm}^3 = (10 \text{ cm})^3 = 1000 \text{ cm}^3$

$\Rightarrow 1 \text{ cm}^3 = 1 \times 10^{-3} \text{ dm}^3 = 1 \times 10^{-3} \text{ L} = 1 \text{ mL}$

\Rightarrow **1 cm³ = 1 mL**

Energy units

- In the SI system: the joule (J)
- $1 \text{ J} = 1 \text{ kg}\cdot\text{m}^2/\text{s}^2$
- (example: kinetic energy = $\frac{1}{2} \times \text{mass} \times \text{velocity}^2$)
- The joule is small, therefore kilojoules (kJ) are used in chemistry.

- The calorie (cal) is an older non SI unit.
- $1 \text{ cal} = 4.184 \text{ J}$
- $1 \text{ cal} = \text{energy required to heat } 1.00 \text{ g of pure water from } 14.5 \text{ }^\circ\text{C to } 15.5 \text{ }^\circ\text{C}.$
- $1 \text{ kcal} = 1000 \text{ cal}$

Measurements - Precision, Accuracy



Poor precision,
poor accuracy



Good precision,
poor accuracy



Good precision,
good accuracy

Precision - How well a series of measurements agree with each other.

Accuracy - How close the measurements are to the true value.

Measurements - Experimental Error



Since the year 2000, the penny weighs 2.35 g

[http://en.wikipedia.org/wiki/Penny_\(Canadian_coin\)](http://en.wikipedia.org/wiki/Penny_(Canadian_coin))

<http://www.coleparmer.ca/catalog>

Penny	Mass (g)	Dev. (g)
1	2.34	0.00
2	2.33	0.01
3	2.35	0.01
4	2.33	0.01
5	2.35	0.01

Averages 2.34 ± 0.01
 The % rel. error = 0.4 %

More **precise** →
 ← More **accurate***

Penny	Mass (g)	Dev. (g)
1	2.3215	0.0014
2	2.3188	0.0013
3	2.3206	0.0005
4	2.3192	0.0009
5	2.3203	0.0002

Averages 2.3201 ± 0.0009
 The % rel. error = 1.27 %

$$\text{The \% relative error} = \frac{|\text{Accepted value} - \text{Observed value}|}{\text{Accepted value}} \times 100\%$$

* So who forgot to zero their balance, causing a **systematic error**?!

Measurements – Standard Deviations

Standard deviation: In a series of N measurements, it is equal to the square root of the sum of the squares of the deviations for each measurement from the average divided by N-1.

Statistical significance in a series of a large number of measurements:

- 68% of the values are expected to be within one standard deviation;
- 95% within two standard deviations

=====
Determination Measure(g) Average-Measure(g) Square of Difference

1	9.990	-0.006	4×10^{-5}
2	9.993	-0.009	8×10^{-5}
3	9.973	0.011	12×10^{-5}
4	9.980	0.004	2×10^{-5}
5	9.982	0.002	4×10^{-6}

Average = 9.984 g

Sum of squares of differences = 26×10^{-5}

Standard Deviation = $\sqrt{(26 \times 10^{-5}/4)} = 0.008$

Mathematics of Chemistry

Exponential or Scientific Notation

$$1234 = 1.234 \times 10^3 \quad (1234 : \text{fixed notation})$$

$$0.01234 = 1.234 \times 10^{-2} \quad (1.234 \times 10^{-2} : \text{exponential notation or scientific notation})$$

Adding or subtracting: First, convert to the same power of 10

$$(1.234 \times 10^{-3}) + (5.623 \times 10^{-2}) = ?$$

$$(0.1234 \times 10^{-2}) + (5.623 \times 10^{-2}) = 5.746 \times 10^{-2}$$

Multiplying: Add the exponents

$$(6.0 \times 10^{23}) \times (2.0 \times 10^{-2}) = (6.0)(2.0) \times 10^{23-2} = 12.0 \times 10^{21} = 1.2 \times 10^{22}$$

Dividing: Subtract the exponents

$$(7.6 \times 10^5) / (1.23 \times 10^2) = [(7.6)/(1.23)] \times 10^{5-2} = 6.18 \times 10^3$$

Measurements - Significant Figures

The amount of mercury taken in a single sample could be specified as 24 ± 0.1 mL indicating the possible error explicitly, or as 24.0 mL, which simply implies that the last digit is uncertain. This particular number is said to have 3 *significant figures*.

Rules for reading and writing experimentally determined numbers:

0.000361

25.00271

16.82600

1230.

12700

Non-zero digits are *always* significant.

Zeros - there are three kinds:

Leading zeros do not count as significant figures.

Captive (imbedded) zeros always count as significant figures.

Trailing zeros are significant *provided the number is written with a decimal point*.

Trailing zeros without a decimal point being present are ambiguous. *Never* write a number which is not an exact number (integer) this way. *Always* use scientific notation to avoid the ambiguity e.g. 1.270×10^4 .

Measurements - Significant Figures

Rules for exact numbers:

These are assumed to have an infinite number of significant figures.

They include:

Integers in formulae, e.g. the volume of a sphere = $\frac{4}{3}\pi r^3$ where the 3 and the 4 are exact numbers.

Defined conversion factors, that is, not experimentally determined, e.g. the US inch is defined as 2.54 cm. The conversion factor between degrees kelvin and celsius *is* experimentally determined and is given to 5 significant figures in the text, 273.15 °.

Calculations - Significant Figures

Rules:

Multiplications and divisions: The result has the same number of significant figures as the *number of significant figures* of the least precise quantity in the calculation.

$$\frac{27.62 \times 36.2738}{1324.1} = 0.7766515 \rightarrow 0.7767$$

Addition and subtraction: The result must have the same number of *decimal places* as the quantity with the fewest.

Example 1:

$$12.11 + 18.0 + 1.013 = 31.123 \rightarrow 31.1$$

Example 2:

$$1.26 \times 10^3 + 37.1 = 1260. + 37.1 = 1297.1 \rightarrow 1300. \rightarrow 1.30 \times 10^3$$

Calculations - Rounding

Rules:

In a series of calculations, *always carry the extra digits through to the final result*. Only then, if you have to report intermediate results, round them, together with the final answer. Rounding intermediate results as you go along introduces *rounding errors*.

When a number is rounded off, the last digit to be retained (the least significant digit) is increased by 1 if the next digit (the first non-significant digit) is greater than, or equal to 5.

Never make the mistake of rounding sequentially from the right: to round 4.348 to two significant figures, go straight to the 4 and round to 4.3. Do not round to 3 significant figures using the 8 first to give 4.35 and then using the 5 to give 4.4!

An Example

An irregularly shaped piece of very dense metal is found lying in the road near a nuclear power station, and is found to weigh 6.823 g. When it is dropped into a small measuring cylinder filled to 14.35 mL with water, the level rises to 14.70 mL.

Is the metal merely harmless tungsten (density* = 19.60 g cm⁻³), or a piece of radioactive plutonium (density* = 19.82 g cm⁻³) that fell off a truck?

To be worked out in class....

$$V = 14.70 \text{ mL} - 14.35 \text{ mL} = 0.35 \text{ mL} = 0.35 \text{ cm}^3$$

$$\text{Density} = m/V = 6.823 \text{ g} / 0.35 \text{ cm}^3 = 19.49428571 \text{ g cm}^{-3}$$

$$\text{Density} = 19 \text{ g cm}^{-3}$$

Is the metal tungsten or plutonium?

* http://www.engineeringtoolbox.com/metal-alloys-densities-d_50.html

Another Example

There are five hydrocarbons (compounds containing carbon & hydrogen) with formula C_6H_{14} , with a different "connectivity" of atoms. All five are liquids at room temperatures, but they have slightly different densities

You have a pure sample of one of these isomers, and you hope to identify it by its density. You measure the volume of your sample using a graduated cylinder, and its mass using an analytical balance.

1. What is the density of the liquid?
2. Can you conclude its identity within the limits of your experimental error?
3. If you make a more accurate volume measurement of 4.93 cm^3 , how does your conclusion change?

Hydrocarbon isomer	Density (g/cm^3)
2,2-dimethylbutane	0.6600
2-methylpentane	0.6532
1-methylpentane	0.6645
hexane	0.6616
3,3-dimethylbutane	0.6485

Data for your pure sample	
Volume	5.0 cm^3
Mass	3.2745 g
Density	
Volume (better)	
Density	

Dimensional Analysis

Getting the units of the result in a calculation

Number in original unit x Conversion factor = Number in new unit

Conversion factor = new unit/original unit = 1

Expressing 1.025 g in kg: $1.025 \text{ g} \times (1 \text{ kg} / 1000 \text{ g}) = 1.025 \times 10^{-3} \text{ kg}$

Problem: The density of sea water is 1.025 g/cm^3 . Convert it to kg/m^3

$$1.025 \text{g.cm}^{-3} \times \frac{1 \text{ kg}}{1000 \text{ g}} \times \frac{(100 \text{ cm})^3}{1 \text{ m}^3} = 1.025 \text{g.cm}^{-3} \times \frac{1 \text{ kg}}{1 \times 10^3 \text{ g}} \times \frac{1 \times 10^6 \text{ cm}^3}{1 \text{ m}^3}$$

$$1.025 \text{g.cm}^{-3} = 1.025 \times 10^3 \text{ kg/m}^3$$