AP Chemistry

Chapter 1 Outline

a) The Study of Chemistry
   i) Matter: the physical material of the universe; has mass and occupies space
   ii) Property: any characteristic that allows us to recognize a particular type of matter and to distinguish it from other types
   iii) Element: basic substance of matter; about 100 different types; can't be broken down into simpler substances [Images]
   iv) Atom: tiny building blocks of matter; each element has its own kind of atom
      (1) Composition: summary of the kinds atoms in a particular type of matter
      (2) Structure: the arrangement of the atoms in a particular type of matter
   v) Molecules: two or more atoms joined in specific arrangements/shapes
   vi) Goal of chemistry: explaining macroscopic behaviors using submicroscopic descriptions

b) [Classifications of matter]
   i) Physical State, aka states of matter
         (a) Gas
         (b) Liquid
         (c) Solid
   ii) Pure substance: matter that has distinct properties, uniform composition from sample to sample
      (1) Elements: contain only 1 type of atom
         (a) 116 known elements
         (b) Chemical symbols arranged in periodic table
      (2) Compounds: contain 2 or more kinds of atoms, but only 1 kind of molecule;
         (a) Can be decomposed into simpler substances by chemical means
         (b) Have different properties from their constituent elements
         (c) Law of Definite Proportions (aka constant composition)—**Joseph Proust** (~1800)—the elemental composition of a pure substance is always the same
   iii) Mixtures: combinations of 2 or more substances in which each substance retains its chemical identity
         (a) May be heterogeneous: composition, properties and appearance vary throughout
         (b) May be homogeneous: uniform throughout; also known as solutions

c) Properties of Matter
   i) Every substance has a unique set of properties.
   ii) Physical properties: can be measured without changing identity or composition of substance
         (a) Color, odor, density, melting point, hardness, etc.
   iii) Chemical properties: describe the way a substance may change (react) to form other substances
   iv) Intensive properties: do not depend on amount of substance
      (1) Temperature, melting point, density

(2) Can be used to identify substances

v) Extensive properties: depend on the quantity/amount of substance
   (1) Mass, volume

vi) Physical changes: physical appearance of substance changes, but not its composition
   (1) Changes of state

vii) Chemical changes (aka chemical reactions): substance transformed into a chemically different substance

viii) Separation of mixtures by taking advantage of the different properties of the components
   (1) Filtration: separation of a solid from a liquid by passing it over a porous medium (filter paper
   (2) Distillation: separation based on different boiling points of substances
   (3) Chromatography: separation based on different abilities of substances to adhere to the surfaces of various solids

d) Units of Measurement
   i) Quantitative Measurements: associated with numbers
   ii) SI units:

Physical Quantity	Name of Unit	Abbreviation
Mass	Kilogram	Kg
Length	Meter	M
Time	Second	s (or sec)
Temperature	Kelvin	K
Amount of substance	Mole	Mol
Electric current	Ampere	A
Luminous intensity	Candela	cd

   (1) Prefixes: See Table 1.5 for the complete list. These are **especially important conversions: 10^6 μg = 10^3 mg = 1 g = 10^{-3} kg**
   (2) K = $^oC + 273.15$
   (3) $^oC = \frac{5}{9}(^oF - 32)$ OR $^oF = \frac{9}{5}(^oC) + 32)$
   (4) Absolute zero: lowest possible temperature
   (5) Common non-SI volume units: mL, cm^3, L, dm^3
      (a) Common devices to measure volume: syringes, burets, pipets, graduated cylinders, volumetric flask
   (6) $Density = \frac{mass}{volume}$
      (a) Densities are temperature dependent; therefore, temperature should be specified when reporting density of a substance

e) Uncertainty in measurement
   i) Exact numbers: defined values (in conversion factors) or counted
   ii) Inexact numbers: numbers obtained by measurement; inexact due to equipment errors or human errors
   iii) Uncertainty always exists for measured quantities.

iv) Precision: measure of how closely individual measurements agree with each other
v) Accuracy: how closely individual measurements agree with correct value
vi) **Significant figures**: Measured quantities are generally reported in such a way that only the last digit is uncertain.
vii) All digits of a measured quantity are significant figures.
- (1) ± notation: one way to express uncertainty, but often not shown (however, it may become relevant in error analysis)
- (2) Counted values have infinite significant figures
- (3) **Significant Figure Rules**:
  - (a) All non-zero digits are significant.
  - (b) Captive zeroes are significant.
  - (c) Leading zeroes are never significant.
  - (d) Trailing zeroes are significant only if the number contains a decimal.
  - (e) In scientific notation, all digits before the exponential term are significant.
  - (f) When performing calculations using measured quantities, the least certain measurement limits the certainty of the calculate quantity.
    - (i) When adding and subtracting, round based on fewest decimal places.
    - (ii) When multiplying and dividing, round based on fewest significant figures.

f) Dimensional Analysis
- i. Use of "conversion factors" with accompanying units to aid in problem solving
  - 1. Ratios, often considered to have infinite significant figures

$$? \text{ s} = 2.0\ \cancel{\text{yr}} \times \frac{365\ \cancel{\text{days}}}{1\ \cancel{\text{yr}}} \times \frac{24\ \cancel{\text{hr}}}{1\ \cancel{\text{day}}} \times \frac{60\ \cancel{\text{min}}}{1\ \cancel{\text{hr}}} \times \frac{60\text{ s}}{1\ \cancel{\text{min}}}$$
$$= 6.3 \times 10^7 \text{ s (to 2 significant figures)}$$