# Ch. 1 – Chemistry: An Introduction

**Chemistry:** The branch of science concerned with the study of matter, its composition, and its interaction. Matter is the "stuff" of which the universe is made of and has both mass (inertia) and volume (occupies space).

**Phenomenon:** Anything which can be directly or indirectly observed by the senses.

**Observation:** A qualitative or quantitative description of a phenomenon.

**Fact:** A repeatable observation.

**Data:** Recorded observations.

**Qualitative Data (Attributes Data):** A recorded observation of a feature, character, or attribute of a phenomenon.

**Quantitative Data (Variables Data):** A recorded measurement of a property of a phenomenon.

**Inference:** Reasoning based on observation and experience.

**Hypothesis:** An inferred possible explanation of an observation which is usually constructed to promote testability.

**Experiment:** A set of observations designed to test a hypothesis or gather additional data.

**Theory:** A conceptual model based on a set of tested hypotheses that explains a range of natural phenomena.

**Law (Natural Law):** A summarized statement of observed natural behavior which is often expressed in mathematical terms.

**Scientific Method:**
1. Observation
2. Hypothesis
3. Experimentation
4. Theory
5. Further Experimentation
6. Modified Theory

Steps 1-3 are continued until a large body of tested hypotheses exist for a given range of phenomena. A theory is generated based on these hypotheses. Steps 5-6 are repeated indefinitely as the theory is continually tested and modified.

# Ch. 2 – Measurements and Calculations

## 2.1 – Scientific Notation

**Decimal Notation:** A system of numeration (counting) that uses a base 10.

**Scientific Notation:** An alternative (usually shorter) way to express a number commonly used to represent very large or very small numbers. The number is expressed as a product of a number with magnitude between 1 and 10 and a power of 10.

**Converting from Decimal Notation to Scientific Notation:**

**Ex 2.1.1:** Convert the following numbers into scientific notation.

**(a)** 20,356,500 = ? $2.03565 \times 10^{7}$

(b) 0.01234 = ? $1.234 \times 10^{-2}$

(c) –515,670,000,000,000 = ? $-5.1567 \times 10^{14}$

(d) –0.0000000000065 = ? $-6.5 \times 10^{11}$

(e) 2.68 $2.68 \times 10^{2}$

**Converting from Scientific Notation to Decimal Notation:**

**Ex 2.1.2:** Convert the following numbers into decimal notation.

(a) $1.176 \times 10^{-6}$ = ? .000001176

(b) $-3.48 \times 10^{5}$ = ? −348,000

(c) $4.8975 \times 10^{-12}$ = ? .0000000000048975

(d) $10^{4}$ = ? 10,000

(e) $-10^{-3}$ = ? −.001

## 2.2 & 2.3 – Metric & SI Units

**Units:** All measurement must be made relative to commonly acceptable standards called ***units***.

**Metric System:** A decimal system of measurement. All units are related to each other by factors of 10 (multiples of 10). A metric unit consists of a prefix and base unit.

metric unit = (prefix)(base unit)

**Metric Base Units:**

| Physical Quantity | Unit Name | Base Symbol |
|---|---|---|
| mass | gram | g |
| length | meter | m |
| volume | liter | L |
| time | second | s |
| temperature | Kelvin | K |

**Metric Fundamental Unit:** A metric unit with no prefix is just a base unit and is sometimes called a ***fundamental unit.***

**SI (Système International):** A comprehensive ***International System*** of measurement based on the metric system. SI is the preferred measurement convention for scientific research.

**Equivalence Relation:** A relation where two members of the same dimensional set (length, area, volume, weight, speed, etc.) are equivalent. A quantitative relationship where two numbers with different units stand for the same measurement.

**Volume Equivalence Relations:** A liter is equivalent to the volume of 1 decimeter cube. A milliliter is equivalent to the volume of 1 centimeter cube.

1 liter (L) = 1 $dm^3$

1 milliliter (mL) = 1 $cm^3$ = 1 cc

*Note: Cubic centimeter ($cm^3$) is also abbreviated as "cc".*

# Common Metric Prefixes

***You are responsible for memorizing all the prefixes from giga to pico!***

| Prefix | Symbol | Meaning | Power of 10 |
|---|---|---|---|
| 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. | $10^{2}$ |
| deca- | da | 10. | $10^{1}$ |
| **no prefix** | *** | 1. | **$10^{0}$** ← **Fundamental Unit** |
| 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 | $10^{-18}$ |

*Note: The symbol "μ" used to designate the prefix micro is the greek letter "mu" and is pronounced like the word "moo".*

**Metric Conversions:** To convert from one metric unit to another requires only the movement of the decimal.

**Ex 2.2.1:** Complete the following conversions.

(a) 2.3 km = 230000 cm

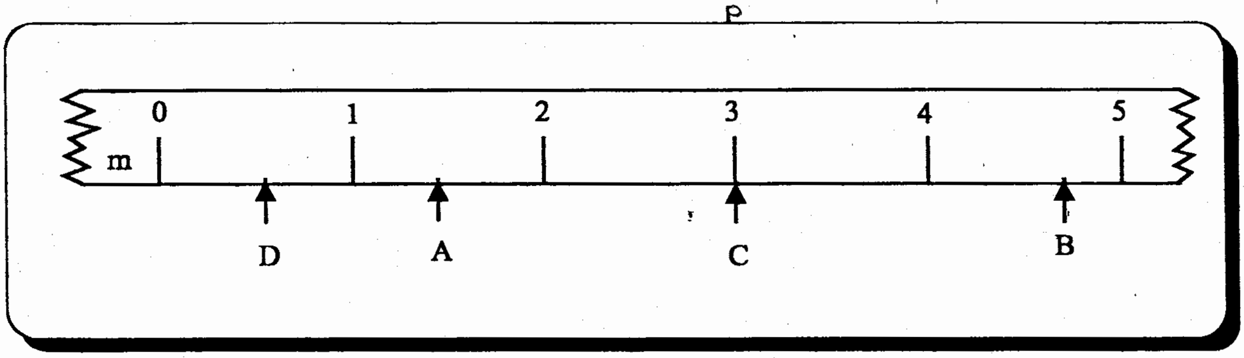
(b) 756 g = 75.6 dag

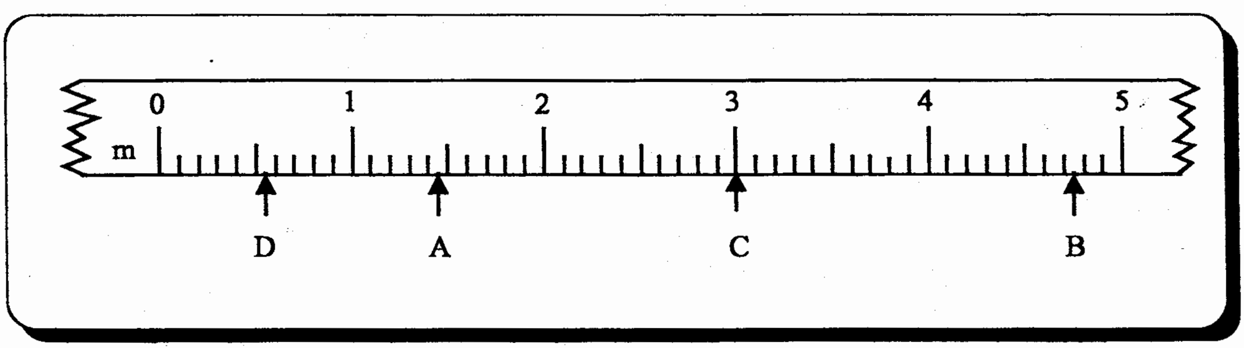
(c) 0.0985 Mm = 98500 m

(d) 54,500,000 pL = 54.5 μL

G = M = k h da xx d c m = μ = n = p

## 2.4 – Uncertainty in Measurement



**Accuracy:** The degree to which a measured value agrees with the true value.

**Precision:** The degree to which a set of measurements have the same value. Precision represents the reproducibility or uncertainty of a measurement and indicates the limitations of the measuring instrument.

**Certain Digit:** A digit whose value is certain since the scale mark for this digit falls to the left of the mark being read.

**Uncertain Digit:** A digit whose value is uncertain since the value of the digit must be guessed or estimated.

**Doubtful Digit:** The first uncertain digit of a measurement. The smaller the magnitude of the doubtful digit (i.e., the farther the digit is to the right relative to the decimal), the greater the precision of the measurement.

**Significant Figures:** A common convention used to indicate the precision of a measuring instrument. The significant figures of a measurement are all the digits known with certainty and the first digit that is uncertain.

**Uncertainty:** The uncertainty of a measurement is $\pm$ 1 of the doubtful digit.

**Range:** The stated limits (minimum and maximum) within which the true value of a measurement will lie. The range of a measurement is determined by adding and subtracting 1 of the doubtful digit to the value of the measurement.

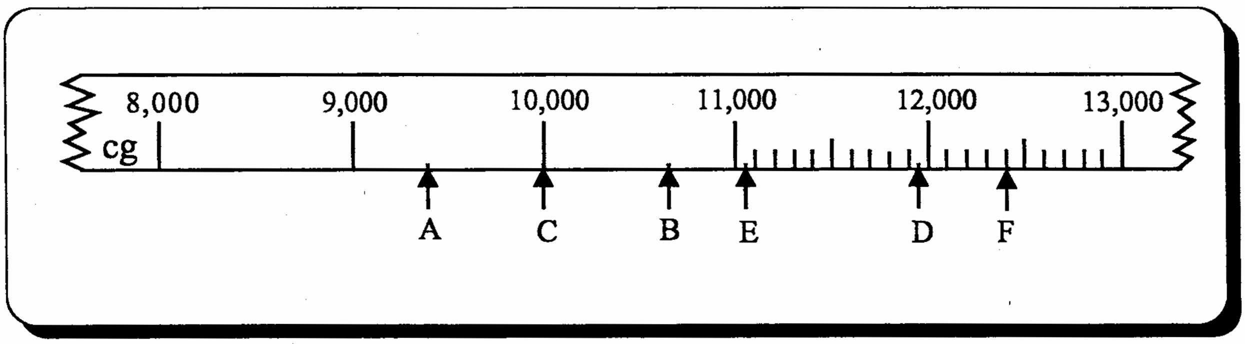
| ***e.g.,*** | Measurement | Uncertainty | Range |
|---|---|---|---|
| | 2,500 | $\pm$ 100 | 2,400 — 2,600 |
| | 25 | $\pm$ 1 | 24 — 26 |
| | 0.025 | $\pm$ 0.001 | 0.024 — 0.026 |

**Reading a Scale:** All measurements contain uncertainty since at least one digit in the reading must be estimated. To ensure that greater precision is not implied than actually exists, measurements are read from a scale so that the measurement value contains only one uncertain digit.

*Note: (1) It is possible to have a measurement value with <u>no</u> certain digits. (2) All measurements from the same scale will have their doubtful digit in the same place.*

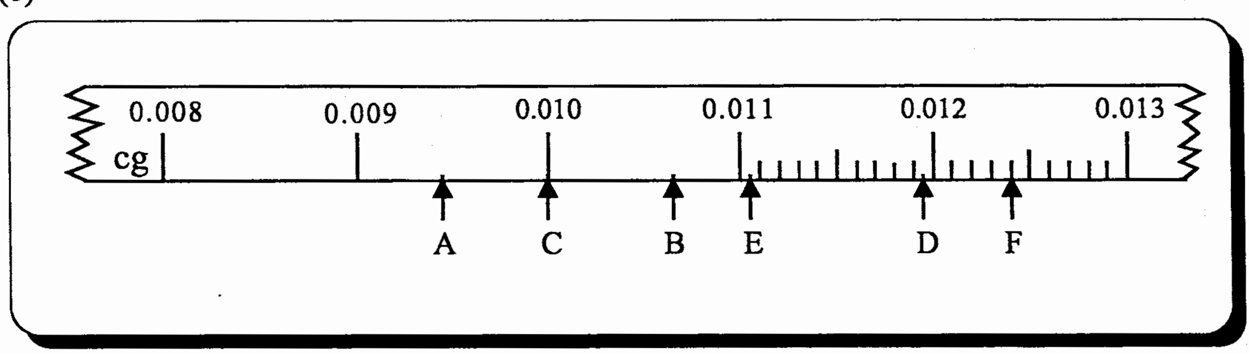
**Ex 2.4.1:** Determine the value, uncertainty, and number of significant figures for each measurement on the following scales. Circle the doubtful digit in the measurement value.

(a)




| Mark | Measurement Value (cg) | Uncertainty | Significant Figures | Mark | Measurement Value (cg) | Uncertainty | Significant Figures |
|---|---|---|---|---|---|---|---|
| A | | | | D | | | |
| B | | | | E | | | |
| C | | | | F | | | |

(b)




| Mark | Measurement Value (cg) | Uncertainty | Significant Figures | Mark | Measurement Value (cg) | Uncertainty | Significant Figures |
|---|---|---|---|---|---|---|---|
| A | | | | D | | | |
| B | | | | E | | | |
| C | | | | F | | | |

## 2.5 – Significant Figures

**Leading Zeros:** Zeros that precede the first non-zero digit.

**Captive Zeros:** Zeros that fall between two non-zero digits.

**Trailing Zeros:** Zeros that follow the last non-zero digit.

**Counting Significant Figures:**

1. Always count non-zero digits.

2. Never count leading zeros.

3. Always count captive zeros.

4. Always count trailing zeros if there is a decimal point shown in the number.

5. Never count trailing zeros if there is no decimal point shown in the number.

*Note: Some books do count trailing zeros in a number with no decimal point as significant. To bypass this possible confusion, scientists and engineers use scientific notation.*

6. If there is a trailing zero with a bar above it, count the trailing zero with the bar and all trailing zeros to the left of it.

*Note: Never use a bar over a digit when a decimal point is shown in the number. The showing of a decimal point in a number overrides the use of a bar over a zero. Never place a bar over a non-zero digit.*

7. Exact numbers (by definition) have infinite significant figures.

8. Scientific notation is treated the same as decimal notation where the number of significant figures is determine from the number (in front) multiplied by the powers of ten.

**Ex 2.5.1:** For each of the following numbers determine the number of significant numbers and circle the doubtful digit.

(a) 12,345.67 (b) 0.0000123

(c) 102.30045 (d) 102,300.

(e) 102,300.0 (f) 0.001200

(g) 12,030,0$\bar{0}$0 (h) $1.00200 \times 10^{-3}$

(i) $0.0012350 \times 10^{5}$ (j) $10.30 \times 10^{-7}$

**Maintaining Significant Figures During Conversions:** When converting from scientific notation to decimal notation and vice versa, or from one unit to another, the number of significant digits in the ending number should be the same as the beginning number.

**Ex 2.5.2:** Change the following numbers into scientific notation.

(a) 41,500,$\bar{0}$00 =

(b) 0.00012400 =

**Ex 2.5.3:** Change the following numbers into decimal notation.

(a) $5.6070 \times 10^{7}$ =

(b) $3.0400 \times 10^{-5}$ =

**Ex 2.5.4:** Complete the following conversions.

(a) 0.1300 km = ____________________ mm

(b) 67,00$\bar{0}$,000 g = ____________________ dg

**Rules for Rounding Off:**

1. If the doubtful digit is followed by 0, 1, 2, 3, or 4, then <u>round down</u>.

2. If the doubtful digit is followed by 6, 7, 8, or 9, then <u>round up</u>.

3. If the doubtful digit is followed by 5 and this 5 is followed by any non-zero digits, then <u>round up</u>.

4. If the doubtful digit is followed by 5 and this 5 is followed by no digits or only zero digits, then <u>round to the closest even number</u>.

*Note: The number in this case lies exactly between the next lower and upper rounded number. By rounding off to the even number, the error due to rounding is averaged out. Half the time the number is rounded up (when doubtful digit is odd) and half the time the number is rounded down (when doubtful digit is even).*

**Ex 2.5.5:** Round the following numbers to the specified number of significant numbers.

(a) 257.52, 4 sig. figs. 257.5

(b) $5.678 \times 10^{-4}$, 3 sig. figs. $5.68 \times 10^{-4}$

(c) 0.004750010, 2 sig. figs. .0050

(d) $5.7983652 \times 10^{12}$, 6 sig. figs. $5.79837 \times 10^{12}$

(e) 26.925, 4 sig. figs. 26.92

(f) 26.925000, 4 sig. figs. 26.92

(g) 26.935, 4 sig. figs. 26.94

(h) 26.935000, 4 sig. figs. 26.94

**Multiplication/Division:** The number of significant figures of the answer is the same as the fewest number of significant figures of the numbers multiplied or divided.

**Ex 2.5.6:** Complete the following calculations ~~with the correct significant figures.~~

(a) $(60{,}450{,}\bar{0}00)(0.0014080) =$ 85114

(b) $(7.0900 \times 10^{-4})/(2{,}879.000) =$ $2.4627 \times 10^{-7}$

(c) $(1{,}000{,}000)(.9999) =$ 999900 = 1 mil

(d) $(89{,}000.)/(1.000 \times 10^{-3}) =$ 89,000,000 or 90,000,000

**Addition/Subtraction:** The doubtful digit of the answer should be in the same digits place as the doubtful digit that lies farthest to the left of the numbers added or subtracted.

**Ex 2.5.7:** Complete the following calculations with the correct significant figures.

(a) $7{,}790.2 + 9{,}6\bar{0}0 =$

(b) $0.0005067 - 3.450 \times 10^{-2} =$

(c) $1{,}000 + 1 =$

(d) $1.00 \times 10^{6} - 35{,}\bar{0}00 =$

## 2.6 – Dimensional Analysis

**Equivalence Relation (Equivalence Statement):** A relation where two members of the same dimensional set (*e.g.*, length, area, volume, weight, speed) are equivalent. A quantitative relationship where two numbers with different units stand for the same measurement.

*Note: Since equivalence relations can always be determined to an accuracy beyond that of any measurement, the numbers in an equivalence relation will be treated as exact numbers.*

***e.g.,*** 1 ft = 12 in
1 kg = 1,000 g

**Unit Ratio (Conversion Factor):** A ratio formed by dividing one side of an equivalence relation by the other. A unit ratio is always equal to the number 1.

***e.g.,*** $\frac{1\text{ ft}}{12\text{ in}} = \frac{12\text{ in}}{12\text{ in}} = 1 \quad \Rightarrow \quad \frac{1\text{ ft}}{12\text{ in}} = 1$

$\frac{1\text{ ft}}{1\text{ ft}} = \frac{12\text{ in}}{1\text{ ft}} = 1 \quad \Rightarrow \quad \frac{12\text{ in}}{1\text{ ft}} = 1$

***e.g.,*** $\frac{1\text{ kg}}{1{,}000\text{ g}} = \frac{1{,}000\text{ g}}{1{,}000\text{ g}} = 1 \quad \Rightarrow \quad \frac{1\text{ kg}}{1{,}000\text{ g}} = 1$

$\frac{1\text{ kg}}{1\text{ kg}} = \frac{1{,}000\text{ g}}{1\text{ kg}} = 1 \quad \Rightarrow \quad \frac{1{,}000\text{ g}}{1\text{ kg}} = 1$

**Dimensional Analysis:** A mathematical technique used extensively in science and engineering for converting units. A measurement is multiplied by one or more unit ratios until the desired unit is obtained.

*Note: Since the numbers in unit ratios are exact numbers, the correct number of significant figures in the answer will be the same number of significant figures as in the starting value.*

### Single-Step Dimensional Analysis

**Ex 2.6.1:** Convert 2.37 centimeters to inches.

2.37 cm = _____ in

available equivalence relation
1 in = 2.54 cm

**Multiple-Step Dimensional Analysis**

**Ex 2.6.2:** Convert 5,650 feet to kilometers.

5,650 ft = _____ km

available equivalence relations
1 ft = 12 in
1 in = 2.54 cm

**Higher Order Dimensional Units:** Units which are taken to a power other than one.

***e.g.,*** Area: $cm^2$, $in^2$, $ft^2$
Volume: $cm^3$, $in^3$, $ft^3$

**Using Dimensional Analysis to Convert Higher Order Dimensional Units:**

**Ex 2.6.3:** Convert 25 square feet to square centimeters.

25 $ft^2$ = _____ $cm^2$

available equivalence relations
1 in = 2.54 cm
1ft = 12 in

**Complex Units:** Complex units are units which are a composite of two or more simple units.

| | | |
|---|---|---|
| ***e.g.,*** | Density: | $g/cm^3$ |
| | Speed: | mi/hr |
| | Acceleration: | $ft/s^2$ |
| | Flowrate: | gal/min |
| | Universal Gas Constant: | $R = 0.08206 \frac{\text{L-atm}}{\text{K-mol}}$ |

**Using Dimensional Analysis to Convert Complex Units:**

**Ex 2.6.4:** Convert 2.17 grams per milliliter to pounds per cubic foot.

2.17 g/mL = _____ $lb/ft^3$

available equivalence relations
1 lb = 453.6 g
1 mL = 1 $cm^3$
1 in = 2.54 cm
1 ft = 12 in

## 2.7 – Temperature

**Temperature:** Temperature is a measure of the kinetic energy (energy due to the motion of atoms and molecules) within a body. The faster the atoms and molecules move, the greater the kinetic energy, hence the higher the temperature.

**Heat Flow:** Heat (thermal energy) transfers from a body of higher temperature to a body of lower temperature. Heat flow will not occur when two bodies are at the same temperature.

**Temperature Scale:** A temperature scale is constructed by fixing two temperature points on a scale usually to common physical phenomena and then dividing the distance between the points into equally spaced intervals.

**Fahrenheit Scale:** Using this scale, water freezes at 32°F and boils at 212°F.

**Celsius Scale:** Using this scale, water freezes at 0°C and boils at 100°C.

**Kelvin Scale (Absolute Scale):** This scale uses the same size degree as the Celsius scale but the zero point has been shifted to the coldest temperature possible in the universe (absolute zero). Using this scale, water freezes at 273.15 K and boils at 373.15 K.

**Absolute Zero:** A hypothetical point determined by extrapolating the theoretical relationship between kinetic energy and temperature. Absolute zero is the temperature at which a body would have no kinetic energy.

**Reading Temperature Units:**

100 °F is read as 100 degrees Fahrenheit.

100 °C is read as 100 degrees Celsius.

100 K is read as 100 Kelvins.

*Note: The Kelvin scale unit symbol does* <u>*not*</u> *use the degree symbol " ° " before the letter K.*

## Temperature Conversions

Celsius → Kelvin: $T_K = T_C + 273.15$

Kelvin → Celsius: $T_C = T_K - 273.15$

*Note: The number 273.15 is treated as an exact number, hence the answer should have the doubtful digit in the same digits place as the starting value.*

**Ex 2.7.1:** Complete the following temperature conversion.

123.4 K = __________ °C

Celsius → Fahrenheit: $T_F = 1.8(T_C) + 32$

Fahrenheit → Celsius: $T_C = \dfrac{(T_F - 32)}{1.8}$

*Note: The numbers 1.8 and 32 are treated as exact numbers, hence the answer should have the same number of significant figures as the starting value.*

**Ex 2.7.2:** Complete the following temperature conversion.

81.78°C = __________ °F

Kelvin → Fahrenheit: K → °C → °F

Fahrenheit → Kelvin: °F → °C → K

*Note: Since this is a two step conversion, significant figures must be carried through from the first step to the second step.*

**Ex 2.7.3:** Complete the following temperature conversion.

–73.2°F = __________ K

## 2.8 – Density

**Density (ρ):** Density is the amount of matter present in a given volume. Density is defined as the mass per unit volume.

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

where: $\rho$ = density
$m$ = mass
$V$ = volume

*Note: The symbol "ρ" used to designate density is the Greek letter "rho" and is pronounced like the word "row".*

**Ex 2.8.1:** Calculate the density of a metal alloy that has a volume of 25.3 mL and a mass of 54 g.

**Ex 2.8.2:** Aluminum has a density of 2.70 g/cm$^3$. Calculate the volume (mL) of a 12 g piece of aluminum.

**Volume Displacement Method:**

1. Partially fill a graduated cylinder with water and record the volume. This volume measurement will be called the initial volume.

2. Drop the object in the water and record the volume. This volume measurement will be called the final volume.

3. The volume of the object can be found by taking the difference between the final and initial volumes.

$$V = V_F - V_I$$

where: $V$ = volume of object
$V_I$ = initial volume
$V_F$ = final volume

**Ex 2.8.3:** A graduated cylinder is initially filled with 25.45 mL of water. A coin having a mass of 0.672 g is dropped into a graduated cylinder causing the water level to rise to 27.63 mL. What is the density of the coin?

**Ex 2.8.4:** A graduated cylinder is initially filled with 58.4 mL of water. If a 43.5 g bolt with a density of 5.736 g/mL is dropped into the graduated cylinder, to what volume will the water rise?

# Ch. 3 – Matter and Energy

**Matter:** The physical material of the universe. Matter is anything that has both <u>mass</u> (inertia) and <u>volume</u> (occupies space).

**<u>States of Matter (Physical States or Phases):</u>**

*Solid:* rigid, finite volume, fixed shape

*Liquid:* fluid, finite volume, same shape as its container

*Gas:* fluid, same volume and shape as its container

**<u>Changes of State:</u>**

| | |
|---|---|
| *Solidification (Freezing):* | liquid → solid |
| *Melting:* | solid → liquid |
| *Vaporization (Evaporation):* | liquid → gas |
| *Condensation:* | gas → liquid |
| *Sublimation:* | solid → gas |
| *Deposition:* | gas → solid |

**Vapor:** The gaseous form of a substance which exists in a liquid or solid state at normal ambient conditions.

**Physical Property:** A characteristic of a substance that can change without the substance becoming a different substance.

**Chemical Property:** A characteristic of a substance related to its ability to react with and change into other substances.

**Physical Change:** A change in the form of a substance, but not in its chemical composition. No chemical bonds are broken or rearranged in a physical change.

**Chemical Change (Chemical Reaction):** The change of a substance into a different substance. Chemical bonds are broken and atoms are rearranged to form one or more different substances.

**<u>Evidence of a Chemical Change:</u>**
color change
bubbling (formation of a gas)
absorption or evolution of heat (change in temperature)
fire or smoke
precipitation (formation of a solid)

**Endothermic Reaction:** A chemical reaction where energy is absorbed.

**Exothermic Reaction:** A chemical reaction where energy is released or evolved.

**Element:** A substance that cannot be decomposed into simpler substances by chemical processes. An element consists of atoms which all have the <u>same number of protons</u>.

**Compound:** A substance formed by the combination of elements in fixed proportions. A compound can be decomposed into elements and/or smaller compounds by chemical processes.

**Molecule:** A chemical structure consisting of two or more atoms. The fundamental chemical unit for some elements and all compounds is the molecule. A molecule of an element is made of only one type of atom while a molecule of a compound must be made of at least two different types of atoms.

**Pure Substance:** Any substance with constant composition. All isolated elements and compounds are pure substances.

**Mixture:** A material of variable composition that contains two or more distinct pure substances in which there is no chemical bonding between them.

**Homogeneous Mixture (Solution):** A mixture that has the same chemical and physical properties throughout all regions of the mixture. The atoms and molecules are uniformly interspersed (mixed together).

*Note: A homogeneous mixture has no visible boundaries between the different components.*

**Heterogeneous Mixture:** A mixture that has different chemical or physical properties in different regions of the mixture. The atoms and molecules are <u>not</u> uniformly interspersed.

*Note: A heterogeneous mixture has one or more visible boundaries between the different components.*

## Physical Separation of Mixtures

**Distillation:** A method of separation based on the differences in volatility (tendency to vaporize).

The mixture is heated until the substance with the lower boiling temperature vaporizes. This vapor can then be drawn off, cooled, and condensed into a separate container. The substance with the higher boiling temperature is left behind as either a solid or liquid.

**Filtration:** A method of separation based on the differences in particle size.

The mixture is poured onto a filter which allows the liquid to pass through retaining the solid.

**Chromatography:** A method of separation based on the differences in solubility (tendency of one substance to dissolve another substance).

The mixture is dissolved in a gas or liquid called the mobile phase. The components are separated as the mobile phase moves over a solid (or viscous liquid) surface called the stationary phase. The component with low solubility spends less time at any particular spot on the stationary phase, thus moving faster and separating away from the component with high solubility.

## Energy and Energy Changes

**Energy:** The measure of a system's ability to do work.

**Heat:** Energy in transit due to a temperature difference. Heat always flows from the higher temperature body to the lower temperature body.

**Calorie [cal]:** The amount of energy required to raise the temperature of one gram of water by one degree Celsius.

**Dietary Calorie [Cal]:** The measurement unit used to measure the amount of energy in food. A dietary calorie is the same unit as a kilocalorie.

1 Cal = 1 kcal = 1,000 cal

**Joule [J]:** The SI unit of energy.

**Specific Heat Capacity (Specific Heat):** The amount of energy required to change one gram of a substance by one degree Celsius (cal/g-°C or J/g-°C).

**Specific Heat Calculations:**

$Q = sm\Delta T$ $\qquad (\Delta T = T_F - T_I)$

$Q = sm(T_F - T_I)$

where: $Q$ = energy absorbed or emitted (cal or J)
$s$ = specific heat capacity (cal/g-°C or J/g-°C)
$m$ = mass (g)
$\Delta T$ = change in temperature (°C)
$T_I$ = initial temperature (°C)
$T_F$ = final temperature (°C)

Sign Conventions: Q will be positive (+) when the system absorbs or gains energy, while Q will be negative (–) when the system emits or loses energy.

**Ex 3.6.1:** Aluminum has a specific heat of 0.8900 J/g-°C. How much heat is required to raise the temperature of a 20.0 g sample from 25°C to 48°C?

**Ex 3.6.2:** What will be the final temperature of 3.58 g of water initially at 32°C, if 250.0 J of heat is absorbed? The specific heat capacity of water is 4.184 J/g-°C.

# Ch. 4 – Chemical Foundations: Elements, Atoms, and Ions

## Chemical Symbols and Formulas

**Element Symbol:** A symbol consisting of one or two letters which designates a specific element. The first letter is always capitalized (upper case) while the second letter is never capitalized (lower case).

**Ex 4.1:** Identify the name of the following elements.

(a) P $\qquad$ (b) Fe

(c) As $\qquad$ (d) Co

**Chemical Formula:** A representation of a compound in which the element symbols designate the type of elements present while the numerical subscripts designate the relative number of atoms. A chemical formula is either an empirical or molecular formula.

**Empirical Formula:** A chemical formula which indicates the simplest whole-number ratio of the elements in a compound.

**Molecular Formula:** A chemical formula which indicates the actual number of each type of atom in a molecule (unit of a compound).

**Ex 4.2:** Identify the number and type of each atom in the following molecules given the molecular formula.

(a) $SO_3$

(b) $C_2H_5O$

(c) $Al_2(SO_4)_3$

**Ex 4.3:** Find the empirical formula given the molecular formula.

| Molecular Formula | Empirical Formula |
|---|---|
| (a) $C_4H_{16}O_8$ | |
| (b) $H_2O$ | |
| (c) $C_6H_{15}O_9N_5$ | |
| (d) $Mg(NO_3)_2$ | |

**History of Modern Atomic Theory**

**1785 – Law of Conservation of Mass:** Antoine Lavoisier (regarded as the "father of modern chemistry") experimentally showed that mass is conserved (neither created nor destroy) in a chemical reaction.

**1799 – Law of Constant Composition (Law of Constant or Definite Proportions):** Joseph Proust experimentally showed that a given compound contains the same proportion of elements by mass no matter how the compound is made.

**1802 – Gay-Lussac's Law:** Joseph Louis Gay-Lussac experimentally showed that when volumes of gases at the same temperature and pressure combine chemically, the volumes of the gaseous reactants form simple ratios to the volumes of the gaseous products.

**1808 – Dalton's Atomic Theory:** John Dalton proposed the first model to successfully explain the observations of the Law of Conservation of Mass, Law of Constant Composition and Gay-Lussac's Law. He accomplished this by reviving the Greek philosopher Democritus' concept that matter is made up of simple building blocks called atoms (from the Greek word *atomos*). Dalton's atomic theory allowed scientists to predict the different possible molecular compositions of compounds made from specific elements.

**Dalton's Atomic Theory:**

1. Elements are made up of tiny particles called atoms.

2. All atoms of a given element are identical.

3. The atoms of a given element are different from those of another element.

4. Atoms of one element combine with atoms of other elements to form compounds. A given compound has the same relative number and types of atoms.

5. Atoms can not be broken into smaller particles by chemical processes. Atoms are never created nor destroyed in chemical reactions, they are simply rearranged into new combinations.

**1811 – Avogadro's Hypothesis:** Amadeo Avogadro interprets Gay-Lussac's Law using Dalton's atomic theory. Avogadro proposes that at the same temperature and pressure, equal volumes of gas contain equal number of particles (atoms and molecules).

**1897 – Discovery of Electron:** J. J. Thomson showed using a cathode-ray tube (Crookes tube) that atoms contain electrons. A cathode-ray tube is a sealed glass tube which contains a gas and two separated metal electrodes. The cathode-ray tube emits a cathode-ray (glowing beam) from the negative electrode (cathode) to the positive electrode (anode) when an electric voltage is applied across the tube. Thomson placed positive and negative electric fields next to the cathode-ray. The ray was deflected toward the positive field and away from the negative field. For this deflection to occur, Thomson postulated that the cathode-ray must be a stream of negatively charged particles (now called electrons). Furthermore, since the ray is emitted from an electrode, the metal element making up the electrode must also contain these negatively charged particles. All elements can be made to emit electrons so all atoms must contain electrons.

**1897 – Plum Pudding Model:** Lord Kelvin (William Thomson, no relation to J. J. Thomson) proposes the "plum pudding" model of an atom. Lord Kelvin hypothesizes that an atom is like a bowl of plum pudding with raisins. The atom can be pictured as a uniform positively charged spherical cloud (the "pudding") randomly embedded with enough negative electrons (the "raisins") to counterbalance the total positive charge of the cloud.

*Note: J. J. Thomson is frequently given credit for the "plum pudding" model but it is now generally understood that Lord Kelvin rightly deserves the credit.*

**1909 – Charge of Electron:** Robert A. Milikan measured the charge of an electron by suspending between two electric plates very small "atomized" drops of oil containing electrons. A fine mist of oil is sprayed into the apparatus. The oil droplets fall through a hole into a chamber with an upper positively charged plate and a lower negatively charged plate. This chamber is bombarded by x-rays which knock electrons free from the gas molecules in the air. These electrons stick to the oil droplets giving them a negative charge. By adjusting the electric field between the two electric plates, the oil droplets are made to suspend in midair. Knowing the value of the electric field and the mass of the droplets, Milikan was able to determine the charge of an electron.

**1911 – Discovery of Positive Nucleus:** Earnest Rutherford experimentally refutes the "plum pudding" model by discovering that the positive charge of an atom is located in the center of the atom. In his experiment, Rutherford passes an α-particle (alpha-particle) beam of positively charged particles through a gold leaf foil. Based on the "plum pudding" model of an atom, Rutherford expected only minor deflections of the α-particles. His hypothesis did not hold true. Instead of passing right through the gold foil, the α-particles sometimes deflected at large angles and even reflected back to the α-particle source. Rutherford could only explain this deflection and reflection pattern by postulating that most of the mass and positive charge of an atom must be contained in a single very dense spot in the atom most logically located in the center of the atom (now called the nucleus).

**1919 – Discovery of Proton:** Rutherford determines that the nucleus of an atom contains positive particles called protons.

**1932 – Discovery of Neutron:** James Chadwick discovers that neutrons (particles with no charge) also exist in the nucleus of an atom.

## Modern Concept of an Atom

**Atomic Nucleus:** The central core of an atom that contains most of its mass. The nucleus is positively charged and consists of one or more protons and neutrons.

**Electron [$e^-$]:** A negatively charged particle which occupies the space around the nucleus of an atom. An electron has a charge of "−1".

**Proton [p or $p^+$]:** A positively charged particle in the atomic nucleus. The mass of a proton is 1836 times greater than the mass of an electron. A proton has the same magnitude of charge as an electron, but its charge is positive instead of negative. A proton has a charge of "+1".

**Neutron [n or $n^0$]:** A particle in the atomic nucleus which has no charge. The mass of a neutron is 1839 times greater than the mass of an electron. A neutron has a charge of "0".

**Nucleon:** A particle which exists in the nucleus of an atom, either a proton or neutron.

**Atomic Structure:** An atom is made up of a nucleus containing protons and neutrons surrounded by a cloud of electrons. An atom consists mostly of space since the nucleus is very small compared to the entire atom. A neutral atom (no net electric charge) has the same number of electrons as protons.

**Atomic Number [$Z$]:** The number of protons in the nucleus of an atom. Each element has a unique atomic number.

**Mass Number [$A$]:** The total number of protons and neutrons in the nucleus of an atom.

mass number = number of protons + number of neutrons

mass number = atomic number + number of neutrons

$$A = Z + N$$

where: $A$ = mass number
$Z$ = atomic number
$N$ = number of neutrons

**Isotopes:** Two or more atoms of the same element that differ in their mass numbers. Isotopes have the same number of protons but a different number of neutrons.

**Isotope Notation:**

$^{A}_{Z}X$

where: $X$ = element symbol
$A$ = mass number
$Z$ = atomic number

**Isotope Name:**

Element name – $A$

where: $A$ = mass number

***e.g.,*** $^{25}_{13}\text{Al}$ = aluminum-25

$^{27}_{13}\text{Al}$ = aluminum-27

$^{28}_{13}\text{Al}$ = aluminum-28

## Periodic Table

**Period:** A horizontal row on the periodic table. Elements in a period have the same number of electron shells.

**Group (Family):** A vertical column of the periodic chart. Elements in a group have similar chemical properties.

**Metals:** All the elements located to the left of the bold "staircase" except for hydrogen. Metals have the following physical properties:

1. Good heat and electrical conductor
2. Malleable (can be hammered into sheets)
3. Ductile (can be extruded into wire)
4. Lustrous shine

**Nonmetals:** All elements located to the right of the bold "staircase" and hydrogen.

**Metalloids (Semimetals):** Elements with a mixture of both metallic and nonmetallic properties. Elements located immediately adjacent to the bold "staircase" except for aluminum.

| | |
|---|---|
| boron | silicon |
| germanium | arsenic |
| antimony | tellurium |
| polonium | astatine |

**Group Symbol Conventions:**

American Chemical Society (ACS): The first two groups on the left are labeled 1A and 2A, and the last six groups on the right are labeled 3A-8A, from left to right.

International Union of Pure and Applied Chemistry (IUPAC): The groups are number 1-18 from left to right.

*Note: The letter "A" after the group number signifies the ACS group symbol convention. No letter "A" after the group number signifies the IUPAC group symbol convention.*

**Main-Group Elements (Representative Elements):** Groups 1A-8A or Groups 1-2, 13-18

**Transition Metals:** Groups 3-12

**Alkali Metals:** Group 1A or Group 1

**Alkaline Earth Metals:** Group 2A or Group 2

**Halogens:** Group 7A or Group 7

**Noble Gases (Inert Gases):** Group 8A or Group 18

**Lanthanides (Lanthanide Series):** Elements 58-71

**Actinides (Actinide Series):** Elements 90-103

**Natural State of Elements:** The physical state of elements at 25°C. All elements naturally exist in the solid state except for the following liquids and gases.

| Liquid State | Gas State |
|---|---|
| mercury | hydrogen |
| bromine | nitrogen |
| | oxygen |
| | fluorine |
| | chlorine |
| | noble gases |

**Diatomic Molecules:** A molecule composed of two atoms.

**Elements Which Exist in Nature as Diatomic Molecules:**

| | | | |
|---|---|---|---|
| hydrogen: | $H_2$ | fluorine: | $F_2$ |
| nitrogen: | $N_2$ | chlorine: | $Cl_2$ |
| oxygen: | $O_2$ | bromine: | $Br_2$ |
| | | iodine: | $I_2$ |

**Ions:** An atom or molecule that has a net positive or negative charge. Ions form by the gain or loss of electrons. Ions are never formed by changing the number of protons in the nucleus of the atom.

**Cation (+):** An ion with a positive charge. A cation is formed when an atom or molecule loses electrons. There are more protons (positive charges) than electrons (negative charges) resulting in a net positive charge. Metals and hydrogen have a tendency to lose electrons forming cations.

**Anion (−):** An ion with a negative charge. An anion is formed when an atom or molecule gains electrons. There are more electrons (negative charges) than protons (positive charges) resulting in a net negative charge. Nonmetals have a tendency to gain electrons forming anions.

*Note: A good mnemonic for remembering that the cation has the positive charge is to notice that the "t" in the word cation resembles a "+" sign.*

**Electrical Conductivity and Ions:** Any substance containing electrons or ions which can move freely will conduct electricity. An electrical current requires either a moving stream of electrons or a moving stream of ions.

***e.g.*,** Copper wire can conduct electricity because a certain number of electrons in each copper atom are free to move from one atom to another.

***e.g.*,** Sodium chloride contains sodium ions ($Na^+$) and chloride ions ($Cl^-$). When sodium chloride is in a solid crystalline form, the ions are held firmly in a lattice structure preventing ion movement and the conduction of electricity. But if the sodium chloride is melted, the sodium and chloride ions are free to move about allowing the liquid sodium chloride to conduct electricity.

***e.g.*,** Pure water will not conduct electricity because of the absence of any free electrons or ions. But when sodium chloride crystals are dissolved in water, the sodium and chloride ions break apart and move freely within the water allowing the sodium chloride solution (salt water) to conduct electricity.

**Ion Charges and the Periodic Table:** Elements in Groups 1A, 2A, 3A, 6A and 7A typically form ions which have the same number of electrons as the nearest noble gas. Metals lose electrons forming cations with a positive charge while nonmetals gain electrons forming anions with a negative charge.

**Ex 4.6:** Predict the ion that the following elements will most likely form.

(a) Na

(b) O

(c) Al

(d) Cl

(e) Ba

**Ionic Compounds:** Chemical compounds containing both positive ions (cations) and negative ions (anions). Since all neutral compounds must have an overall net charge of zero, the total positive charge from all the cations must be equal to the total negative charge from all the anions.

**Ex 4.7:** Write the chemical formulas for the ionic compounds formed from the following sets of ions.

(a) $K^+$, $I^-$ KI

(b) $Ca^{2+}$, $O^{2-}$ CaO

(c) $Sr^{2+}$, $F^-$ $SrF_2$

(d) $Ba^{2+}$, $N^{3-}$ $Ba_3N_2$

(e) $Pb^{4+}$, $S^{2-}$ $PbS_2$

55 - 76 all
79 - 84 all

## Ex 4.5: Complete the following table.

| | Isotope Symbol | # Protons | # Neutrons | # Electrons | Atomic Number | Mass Number | Charge |
|---|---|---|---|---|---|---|---|
| (a) | $^{53}_{24}Cr^{3+}$ | 24 | 29 | 21 | 24 | 53 | +3 |
| (b) | $^{80}_{35}Br^{-}$ | 35 | 45 | 36 | 35 | 80 | -1 |
| (c) | $^{87}_{38}Sr$ | 38 | 49 | 38 | 38 | 87 | 0 |
| (d) | $^{32}_{15}P^{+3}$ | 15 | 17 | 12 | 15 | 32 | +3 |
| (e) | $^{105}_{46}Pd^{+1}$ | 46 | 59 | 45 | 46 | 105 | +1 |
| (f) | $^{77}_{34}Se^{-2}$ | 34 | 43 | 36 | 34 | 77 | −2 |
| (g) | $^{96}_{45}Mo$ | 42 | 54 | 42 | 42 | 96 | 0 |
| (h) | $^{127}_{53}I^{-1}$ | 53 | 74 | 54 | 53 | 127 | -1 |
| (i) | $^{138}_{56}B^{+2}$ | 56 | 82 | 54 | 56 | 138 | +2 |
| (j) | $^{51}_{23}V^{+3}$ | 23 | 28 | 20 | 23 | 51 | +3 |
| (k) | $^{29}_{14}Si^{-}$ | 14 | 13 | 15 | 14 | 29 | −1 |

[redacted]

Chapter 5 Chem Notes

11/11/03
Chemistry

5.1 -

binary compounds: compounds formed out of two elements

Two types of binary compounds:

1. Metal + Nonmetal (ie table salt)
2. Nonmetal + Nonmetal

5.2 -

binary ionic compound: contain a positive ion (cation), which is always written first, and an anion.

TYPE I COMPOUNDS: the present metal forms only one type of cation.

TYPE II COMPOUNDS: the present metal can form 2+ cations with different charges.

COMMON SIMPLE IONS

| cation | name | anion | name |
|---|---|---|---|
| $H^+$ | hydrogen | $H^-$ | hydride |
| $Li^+$ | lithium | $F^-$ | flouride |
| $Na^+$ | sodium | $Cl^-$ | chloride |
| $K^+$ | potassium | $Br^-$ | bromide |
| $Cs^+$ | cesium | $I^-$ | iodide |
| $Be^{2+}$ | beryllium | $O^{2-}$ | oxide |
| $Mg^{2+}$ | magnesium | $S^{2-}$ | sulfide |
| $Ca^{2+}$ | calcium | | |
| $Ba^{2+}$ | barium | | |
| $Al^{3+}$ | aluminum | | |
| $Ag^+$ | silver | | |

# Chemistry Quiz #4

## Chapter 5

Name: [redacted] Date: 11/13/2003 Period: 6

Write the systematic name for each of the following compounds.

1. $Ag_3AsO_3$ Silver Arsenite

2. $HNO_3(aq)$ Nitric Acid

3. $PCl_3$ Phosphorus Trichloride

4. $Mn_2(S_2O_3)_3$ Manganese (III) Thiosulfate

5. $Ba(BrO_3)_2 \bullet 4H_2O$ Barium Bromate Tetrahydrate

6. $HI(aq)$ Hydroiodic Acid

7. $Ni(CrO_4)_2$ Nickel (IV) Chromate

8. $B_2Cl_6$ Diboron Hexachloride

9. $H_3N$ Hydrogen Nitrade

10. $H_2C_2O_4\ (aq)$ ~~Oa~~ Oxalic Acid

Write the chemical formula for each of the following compounds.

11. Ammonium Hypophosphate Hexahydrate $(NH_4)_2PO_2 \cdot 6H_2O$

12. Auric Cyanide $Au(CN)_3$

13. Calcium Hydride $CaH_2$

14. Hypofluorous Acid $HFO(aq)$

15. Trinitrogen Octachloride $N_3Cl_8$

16. Cadmium(IV) Peroxide $Cd(O_2)_2$

17. Permanganic Acid $HMnO_4(aq)$

18. Disulfur Decafluoride $S_2F_{10}$

19. Hydrophosphoric Acid $H_3P(aq)$

20. Mecurous Acetate $Hg_2(C_2H_3O_2)$