

# Chem\_102\_Notes

## chem\_102-01.19.2023-notes: Sig Figs, Density, Temperature Conversion, Chemistry Concepts

### Sig Figs

#### Example 1

$$\frac{[\text{floor}(\lceil (143.7 - 121) \times 2.06 \rceil)]}{0.600} \approx 79$$

#### Example 2

$$(182.6 - 143.0 + 3.121 - 3.2) = 39.521 \approx 39.5 \quad (> 3 \text{ Sig Figs})$$

#### Example 3

$$\frac{(182.6 - 143.0 + 3.121 - 3.2) \times 4.0}{1.600 \times 0.500} = \frac{39.5 \times 4.0}{1.600 \times 0.500} = 197.5 \approx 2.0 \times 10^2 \quad (> 2 \text{ Sig Figs})$$

### Relative Density (Specific Gravity)

$$\text{SG} = \frac{\text{Density}}{\text{Water Density}}$$

Caution: Distance units in density calculations must be considered.

### Unit Conversion

$$3.51 \text{ g/cm}^3 > \text{kg/dm}^3$$
$$\left(\frac{3.51}{1}\right) \times \left(\frac{1 \text{ kg}}{1000 \text{ g}}\right) \times \left(\frac{(10 \text{ cm})^3}{1 \text{ dm}^3}\right) = 3.51 \text{ kg/dm}^3$$

### Temperature Conversion

$$F = \left(\frac{C \times 9}{5}\right) + 32$$
$$-40^\circ\text{C} = -40^\circ\text{F}$$

## Chemistry Concepts

### Levels of Representation

- Macroscopic
- Microscopic
- Symbolic

### Matter

- Mass
- Volume

## Composition

- Quantitative
- Qualitative

## Properties

- Physical
- Chemical

## Element

Composed of the same type of atoms.

## Element Naming

- Polonium - Named after Poland
- Francium, Gallium - Named after France
- Nihonium - Named after Japan
- Germanium - Named for Germany
- Lecoq de Boisbaudran named Gallium after France and himself.

## The Mole

The amount of substance containing the same number of discrete entities as the number of atoms in a sample of pure Carbon-12 weighing exactly 12 grams. One mole is  $6.02214179 \times 10^{23} \text{ mol}^{-1}$ .

# chem\_102-01.24.2023-notes: Atom, Dalton's Atomic Theory, Experiments, Atomic Components, Isotopes, Formulas, The Mole

## Atom Term Derivation

### Classical Atomic Theorists

- Dalton's Atomic Theory
  - Matter composed of indivisible particles (Standard Model of Subatomic Particles)
  - Elements are composed of identical atoms with a characteristic mass, the same for all atoms of that element (Isotopes)
  - Atoms of one element differ in properties from atoms of other elements
  - A compound consists of atoms of 2 or more elements in a whole number relation
  - Atoms cannot be created nor destroyed in a chemical reaction, merely rearranged into different substances (Daltonian Law of Mass Conservation, Nuclear Reactions - Fusion/Fission Cores)
- Dalton's Theory and Joseph Proust Experiments: Law of Multiple Proportions
  - $\frac{\text{CuCl}}{\text{CuCl}_2} : \frac{63.55}{35.45} \text{ grams vs } \frac{63.55}{70.9} \text{ grams}$
  - $\frac{1.116 \text{ g Cu}}{1 \text{ g Cl}} \div \frac{0.558 \text{ g Cu}}{1 \text{ g Cl}} = \frac{2}{1} \text{ Cu/Cl}$

## Experiments

- Cathode Ray Tube by J.J. Thompson
  - Discovery of the negatively charged Electron ( $\beta^-$ ) and Established prediction of Mass:Charge Ratio
  - Magnetic deflection of beam  $\Rightarrow$  not light since light does not deflect through magnetic fields ( $\Phi$  or  $\beta$ )
- Oil-Drop Experiment by Robert A. Millikan and Harvey Fletcher
  - Calculation of Exact Electron Charge
  - Atomized oil drops negatively ionized via x-ray suspended by parallel metal plates via capacitor
- Ernest Rutherford Gold Foil Experiment
  - Discovery of Nucleus and the Proton
- Becquerel, Curie, and Chadwick Radioactivity Experiment
  - The Neutron and Isotopes: Hydrogen of Mass 1, 2, and 3  $\gg$  Protium, Deuterium, and Tritium

## Atomic Components

- The Scale of Atom in all Models
  - The Atom, The Sub-atomic Particles, and The "Void"
- The Atomic Mass Unit (amu)
  - Defined as  $\frac{1}{12}$  of Carbon-12
  - $\approx 1.662 \times 10^{-27}$
  - Roughly the mass of 1 sample of Protium (Hydrogen-1)
- Absolute Charge of an Electron =  $1.602176634 \times 10^{-19}$  C

## Isotopes

- Atomic Number ( $Z$ ) = # Number of Protons
- Mass Number ( $A$ ) = # Protons + # Neutrons
  - Carbon-12 has 6 Protons, 6 Neutrons, and 6 Electrons
- Isotope: Each of two or more forms of the same element that contains equal numbers of protons but different numbers of neutrons in their nuclei
  - Differ in relative atomic mass but not in chemical properties
- Average Atomic Mass and % Natural Abundance:
  - C-12 = 98.99, C-13 = 1.11, C-14  $\approx$  trace near 0
  - $\Sigma(i) = (\text{frac. abund.} \times \text{iso. mass})$
  - $\frac{(12 \times 98.99) + (13 \times 1.11) + (14 \times 0)}{100} \approx 12.01 \text{ g/mol}$

## Formulas

- Chemical Formula: The representation of a molecule that uses chemical symbols to indicate the types of atoms followed by subscripts to show the number of atoms of each type in the molecule. A subscript is used only when more than one atom of a given type is present.
- Structural Formula (Lewis Structure): A formula that shows the spatial arrangement of atoms in the molecule of a compound.
- Empirical Formula: A formula giving the simplest/lowest proportions of the elements present in a compound but not the actual numbers or arrangement of atoms.

## The Mole

- The amount of substance that contains as many elementary entities as there are atoms in exactly 12 g of carbon-12.
- 1 mole =  $6.02214129 \times 10^{23}$  particles
- Example: 39 g Sucrose (Glucose  $C_6H_{12}O_6$ ) is how many moles?

$$\frac{39 \text{ g}}{1} \times \frac{1 \text{ mol Sucrose}}{(12 \times 12) + (22 \times 1) + (11 \times 16)} \approx \frac{39 \text{ g}}{342 \text{ g/mol}} \approx 0.114 \text{ mol Sucrose}$$

- Example: The same moles of Glucose is how much mass? (Glucose is  $C_6H_{12}O_6$ )

$$0.114 \text{ mol } (C_6H_{12}O_6) \times \frac{(12 \times 6 + 12 \times 1 + 16 \times 6)}{1} \approx 20.5 \text{ g } C_6H_{12}O_6$$

- Example: How many moles of Carbon atoms are infused in 39g of Glucose ( $C_6H_{12}O_6$ )?

$$39 \text{ g Glucose} \times \frac{1 \text{ mol}}{180 \text{ g Glucose}} \times \frac{6 \text{ Carbon atoms}}{1 \text{ Glucose}} = 1.30 \text{ moles of Carbon}$$

## chem\_102-01.31.2023-notes: Electromagnetic Radiation, Quantum Mechanics, Atomic Structure

### Electromagnetic Radiation

$$\lambda \nu = c$$

$\lambda$  = wavelength

$\nu$  = (Greek letter 'mu') frequency of wave

$c$  = speed of light (electromagnetic wave)

$$c \approx 2.998 \times 10^8 \approx 3 \times 10^8$$

### Quantum Mechanics

Max Planck observed that energy can be gained or lost only in whole-number multiples of  $h\nu$ , where  $h$  is Planck's constant ( $6.626 \times 10^{-34}$  J·s).

$$\Delta E = n \cdot h \cdot \nu, \quad \text{where } n \text{ is any natural number } (\mathbb{Z} > 0)$$

Longer & shorter wave  $\rightarrow$  high  $\lambda$ , low  $\nu$ , low  $\Delta E$  e.g., Absorption jump from  $n = 4$  to  $n = 5$

$$E_{\text{photon}} = h\nu = \frac{hc}{\lambda}$$

### The Photoelectric Effect

- The phenomenon whereby electrons are emitted from the surface of a metal when light strikes it
- When  $\nu < \nu_0$ , no electrons are emitted, regardless of the intensity of light
- When  $\nu > \nu_0$ , the number of electrons increases with the intensity of light
- When  $\nu > \nu_0$ , the kinetic energy of emitted electrons increases linearly with the frequency of the light
- Minimum energy required to remove an electron =  $E(0) = h\nu(0)$
- When  $\nu > \nu(0)$ , the excess energy that is required to remove the electron is given as kinetic energy ( $KE$ ):

$$KE(e-) = \frac{1}{2}mv^2 = h\nu - h\nu(0)$$

- Greater intensity of light means that more photons are available to release electrons, which gave rise to the equation:  $E = mc^2$

## Wave-Particle Duality

- The phenomenon whereby electromagnetic radiation (and all matter) exhibits wave properties and particulate properties
- De Broglie's equation allows for the calculation of the wavelength of a particle

$$\lambda = \frac{h}{mv}$$

## Diffraction

- It is the result of light getting scattered from a regular array of points or lines
- This scattered radiation produces a diffraction pattern on bright and dark areas
- Scattered light can:
  - Interfere constructively and produce a bright area
  - Interfere destructively to produce a dark spot
- This phenomenon occurs best when the spacing between scattering points is similar to the wavelength of the diffracted wave

## The Atomic Spectrum of Hydrogen

### Hydrogen Emission Spectrum

- Continuous spectrum occurs when white light is passed through a prism, contains all the wavelengths of visible light
- Hydrogen emission spectrum is called a line spectrum, displays only a few lines, each line corresponding to discrete wavelengths, indicates that the energy of the electron on the hydrogen atom is quantized
- Change in energy from a high to lower level of a given wavelength can be calculated by:  $\Delta E = h\nu = \frac{hc}{\lambda}$

### The Bohr Model

- Quantum Model for the Hydrogen Atom
- Electron in a hydrogen atom moves around the nucleus only in certain allowed circular orbits
- The expression for the energy levels available to the electron in the hydrogen atom can be expressed as:

$$-2.178 \times 10^{-18} \text{ J} \cdot \frac{Z^2}{n^2}$$

- When the hydrogen atom returns to its lowest possible energy state, it is called the ground state
- When the electron falls from  $n = 6$  to  $n = 1$ ,  $\Delta E$  can be computed by:  $\Delta E = E(f) - E(0)$
- The model correctly fits the quantized energy levels of the hydrogen atom, postulates only certain allowed circular orbits for the electron
- As the electron becomes more tightly bound, its energy becomes more negative relative to the free electron
- The free electron is at infinite distance from the nucleus, as the electron is brought closer to the nucleus, energy is released from the system

## The Quantum Mechanical Model of the Atom

- Erwin Schrödinger and Quantum Mechanics
- Standing waves are stationary waves that do not travel along any length; only certain orbits have a circumference into which whole number wavelengths of standing electron waves will fit.
- Other waves produce destructive interference of the standing electron wave.
- The mathematical representation for a standing wave is:  $H\hat{\psi} = E\psi$ , where  $\psi$  represents the wave equation, which is a function of the coordinates of the electron's position in 3-dimensional space.  $H\hat{\psi}$  represents an operator.
- A specific wave function is defined as an orbital.

## Heisenberg Uncertainty Principle

- There is a fundamental limitation to just how precisely we can know both the position and momentum of a particle at a given time
- Mathematically, this principle can be represented by:  $\Delta x \cdot \Delta(mv) \geq \frac{h}{4\pi}$ , where  $\Delta x$  is the uncertainty in a particle's position,  $\Delta(mv)$  is the uncertainty in a particle's momentum, and  $h$  is Planck's constant

## Physical Meaning of the Wave Function

- The square of the function, represented as a probability distribution, indicates the probability of finding an electron near a particular point in space
- The intensity of color is used to indicate the probability value near a given point in space, the more time the electron visits a particular point, the darker the negative becomes
- This diagram is known as an electron density map
- A radial probability distribution graph plots the total probability of finding an electron in each spherical shell versus the distance from the nucleus
- Probability of finding an electron at a particular position is greatest near the nucleus, volume of the spherical shell increases with distance from the nucleus
- The size of the 1s orbital can be stated as the radius of the sphere that encloses 90

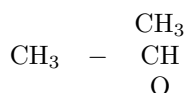
## Quantum Numbers

3rd period can be HyperValent - can go beyond 3 Octet

Will behave like their same period relatives but of 1st or 2nd period, e.g., N likes triple bonds, so Phosphorous bonds in

Empirical Structure, Lewis Structure, Skeletal Structure, Resonance Structures, Formal Charge Calculation, Memorize E

Treat Double/Triple Bonds as Denser Single Bond Protrusions for General Chemistry 1 Purposes, Solid Triangle Out of



Single-bonded Carbons are tetrahedral, oxygenated carbon is trigonal planar

## Continuation

- Memorize Bond Angles and Examples, Molecular Structure and Polarity, VSEPR Theory, Vector Sum of Partial Charge  $\delta$  and Molecular Polarity, Dipole Moments ( $\mu = Qr$ ,  $\mu = 0 \Rightarrow \text{NP}$ ,  $\mu \neq 0 \Rightarrow \text{P}$ )
- C+Cl(4) vs CH+Cl(3) Polar Moments, Analysis of Organic Molecules, Failure of Pure Orbital Model  $\Rightarrow$  Linear Combination of Atomic Orbitals, Hybridization and Hybrid Orbitals, Double:  $sp^2 = \sigma, \pi$ , Single:  $sp = \sigma, 2\pi$ , Butadiene ( $C_6H_6$ ): each C-H bond is  $\sigma$ , C-C is  $\sigma/\pi$ , and C=C is  $\sigma, \pi$
- **Organic Chemistry Notation:** Unless specified, the notation of carbon bonds follows standard conventions.
- **Assignment of Hybridization Involving Resonance:** Understanding hybridization is crucial, especially in cases involving resonance structures.
- **Scientific Experiments and Respective Discoveries:** Familiarize yourself with key experiments in chemistry and their discoveries.
- **Hydrogen Absorption and Emission Spectrum:** Study the absorption and emission spectrum of hydrogen, essential for understanding atomic structure.
- **Ionization Energy vs Electron Affinity vs Atomic Radius & Trends:** Explore the relationships and trends between ionization energy, electron affinity, and atomic radius.
- **Quantum Number Sets:** Understand the sets of quantum numbers (principal, azimuthal, magnetic, and spin) that describe an electron's state.
- **Angular Momentum x, y, z Orientations:** Learn how angular momentum components ( $L_x, L_y, L_z$ ) are used to describe the orientation of an electron's angular momentum vector in space.
- **Orbital vs Suborbital:** Differentiate between orbitals (regions where an electron is likely to be found) and suborbitals (sets of orbitals with the same energy level and shape).
- **Pauli vs Aufbau vs Hund's Rules:** Adhere to the Pauli exclusion principle, Aufbau principle, and Hund's rule when filling electron orbitals. Never break Pauli's principle, and sometimes Aufbau is broken, while Hund's rule is unusual.
- **Switching Electron Spin "Boxes":** It's acceptable to switch up electron spin "boxes" if the Magnetic Quantum Number Convention is not specified (e.g., no -2, -1, 0, 1, 2).
- **Unnecessary Spin Multiplicity:** Understand that unnecessary spin multiplicity increases electron potential energy (e.g.,  $\Delta E(\uparrow\uparrow\downarrow) > \Delta E(\downarrow\downarrow\downarrow)$ ).
- **Wave Equation:** The wave equation  $\lambda = \frac{h}{\text{particle mass} \cdot \text{particle velocity}}$  highlights the relationship between wavelength ( $\lambda$ ), Planck's constant ( $h$ ), momentum, and wave properties.

## chem\_102-02.23.2023-notes: Intermolecular Forces

In general, Intermolecular Forces are weaker than Intramolecular Forces.

- Intramolecular Forces - keep a molecule intact.
- Intermolecular Forces - Hold multiple molecules together and determine many of a substance's properties.

**Intermolecular Forces (IMF)** are essentially understood as Inter Entity Forces.

- Forces of attraction that may exist between any two chemical substances due to electrostatic phenomena.

A permanent/temporary Dipole may form due to the positioning of the Electron Lone Pairs in a certain chemical species and a nearby Ion.

#### **Pentane Isomers?**

Ionic Bond ; Covalent bond ; Metallic Bonds

#### **Nonbonding (Intermolecular) Rankings**

- **Ion-Dipole**

- Formed by Ion Charge and Dipole Charge.
- $\Delta E = 40 - 600 \text{ kJ}$ .

- **H-Bond**

- Formed by Polar bond to H and Dipole Charge (High EN of N, O, F).
- $\Delta E = 10 - 40 \text{ kJ}$ .

- **Dipole-dipole**

- Formed between Dipole Charges.
- $\Delta E = 5 - 25 \text{ kJ}$ .

- **Ion-induced dipole**

- Formed between Ion charge-polarizable electron cloud.
- $\Delta E = 3 - 15 \text{ kJ}$ .

- **Dispersion (London)**

- Formed between Polarizable electron cloud.
- $\Delta E = 0.05 - 40 \text{ kJ}$ .
- Generally assumed to be present between all molecules.
- Bigger unit sizes have stronger bonds.

The strength of ONLY LDF's increases with increased surface area. Everything else is better when smaller.

Hydrogen Bonding must have one Hydrogen bound to H/O/N and the other to F/O/N (**Bond-Network Density**).

#### **Examples:**

- $H_2O$  vs  $KI(aq)$  ; Hydrogen Bonds ; Ionic Bonding  $\therefore KI(aq)$ .
- $HF$  vs  $HI$  ; Fluorine is more electronegative than Iodine, so  $\Delta\chi(HF) > \Delta\chi(HI) \therefore HF$ .
- $Cl_2$  vs  $Br_2$  ; both have LDFs, but Atomic Radius of Br is larger than that of Cl ;  $LDF(Br_2) > LDF(Cl_2) \therefore Br_2$ .

#### **Example: Vapor Pressure Ranking?**

- *Ethanol* vs *EthyleneGlycol* vs *DiethylEther* vs *Water*
- H-bond Polar vs H-bond Polar vs LDF NonPolar vs H-bond, Polar
- Ethylene Glycol has more Hydrogens than Ethanol
- $\therefore$  Bond Strength:  $C_2H_6O_2 > H_2O > C_2H_6O > (C_2H_5)_2O$
- $\therefore$  Vapor Pressure is in reversed order!

#### **Miscellaneous**

- Average: 59, Max: 97.
- Give you a bike, you can ride any version of a bike. That is not a problem; the challenge of which is to test your understanding of the concept.
- No Curve: Mid Term Grade 95.37%, 205.99/216.00 ; Mid Term Exam 01 Mid Term Grade 96/120 =  $0.8 = 80\%$ .

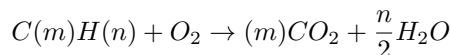


# Quantitative Chemical Analysis

**Percent Composition:** the % mass of each element in a compound

$$\%H = \left( \frac{\text{mass of H in compound}}{\text{mass of compound}} \right) \times 100$$

**Combustion Analysis:**



- Carbon # =  $m$
- $CO_2$  # =  $m$
- Hydrogen # =  $n$
- $H_2O$  # =  $\frac{n}{2}$

$H_2O$  absorber e.g. ???

$CO_2$  absorber e.g. ???

BCA Tables - before, change, and after tables?

Empirical formula - Simplest Ratio:  $CH$

Determining an Empirical Formula from Percent Composition

Balanced Equation  $\rightarrow$  Simplest/Lowest Integer Equation

The five basic types of chemical reactions are: combination, decomposition, single-replacement, double-replacement, and combustion

- $A + B \rightarrow AB$  (Combination)
- $XY \rightarrow X + Y$  (Decomposition)
- $AB + X \rightarrow AX + B$  (Single-Replacement)
- $A^+B^- + X^+Y^- \rightarrow A^+Y^- + X^-B^-$  (Double-Replacement)
- $A + O_2 \rightarrow AO$  (Combustion)

**Oxidation-Reduction (Redox) Reaction:** A type of chemical reaction that involves a transfer of electrons between two species.

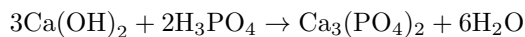
- **Redox:** A type of chemical reaction in which the oxidation states of substrate change.
- **Oxidation:** The loss of electrons or an increase in the oxidation state.
- **Reduction:** The gain of electrons or a decrease in the oxidation state.
- Mnemonics: "OIL RIG" (**O**xidation is **L**oss, **R**eduction is **G**ain), "LEO the lion says GER [grr]" (**L**oss of **E**lectrons is **O**xidation, **G**ain of **E**lectrons is **R**eduction).

**Energy on the Reactants side is Endergonic:** energetically costly reactions which need heat.  
**Energy on the Products side is Exergonic:** energetically favorable reactions which produce heat.

Volumes of Pure Substances  $\Rightarrow$  Masses of Substances  $\Rightarrow$  Moles of Substances  $\Rightarrow$  Balanced Chemical Equation

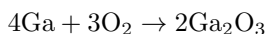
Examples:

- 1.36 mol of (Phosphoric Acid) needs to react with how many moles of  $Ca(OH)_2$  (Calcium Hydroxide)?



$$\therefore 1.36 \times \left( \frac{3}{2} \right) = 2.04 \text{ mol}$$

- What mass of Gallium Oxide can be prepared from 29.0 g of Gallium metal oxidation?

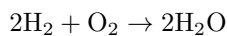


$$\therefore \left( \frac{29.0}{69.723} \right) \times \left( \frac{2}{4} \right) \text{ mol} \times 187.433 \text{ g/mol} = 39.0 \text{ g}$$

## Limiting and Excess Reactants & Cross-check Method for Reagents

### Examples:

- What is the limiting reactant when 5.0 g of  $H_2$  reacts with 10.0 g of  $O_2$  to form water?



$$\frac{5 \text{ g}}{2 \text{ g/mol}} = 2.5 \text{ mol } H_2$$

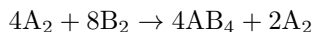
$$2.5 \text{ mol } H_2 \rightarrow \text{need } 1.25 \text{ mol } O_2$$

$$\frac{10 \text{ g}}{32 \text{ g/mol}} = 0.3125 \text{ mol } O_2$$

$$0.3125 \text{ mol } O_2 \rightarrow \text{need } 0.625 \text{ mol } H_2$$

$\therefore O_2$  is the limiting reactant

- 1.0 mol  $A_2$  and 1.0 mol  $B_2 \rightarrow \#$  product?



$$\text{Need } A_2 : B_2 = 2 : 1$$

$$\text{If prod } A_2 = \text{excess? } \therefore (\text{prod } \#) = \left( \frac{\#B_2}{2} \right) = 0.5 \text{ mol } \# \text{ product}$$

- $C, H, \&O > 20.00$  g combustion of a compound  $\rightarrow 44.04$  g  $CO_2$  and 14.32 g of  $H_2O$ . 6.40 g of this 20.00 g is Oxygen

### Continue the Notes:

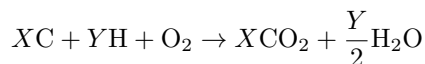
## Limiting and Excess Reactants & Cross-check Method for Reagents

### Examples (Continued):

- Combustion Analysis Steps

$C, H, \&O > 20.00$  g combustion of a compound  $\rightarrow 44.04$  g  $CO_2$  and 14.32 g of  $H_2O$ . 6.40 g of this 20.00 g is Oxygen

Molecular formula? Molar Mass is 200 g/mol



$$\frac{44.04 \text{ g } CO_2}{44 \text{ g/mol}} = \frac{14.32 \text{ g } H_2O}{18 \text{ g/mol}}$$

$$\frac{1}{5} = \frac{0.8}{4}$$

$$XC_{10}H_{16}O_4 = 12 \times 10 + 1 \times 16 + 16 \times 4 = 120 + 16 + 64 = 200$$

$$\frac{6.4}{20.00} = 0.32 = 32\%; \left( \frac{16 \times 4}{200} \right) = \frac{64}{200} = \frac{32}{100} = 0.32 = 32\%$$

$$\frac{44.04 \text{ g } CO_2}{20 \text{ g sample}} \times \frac{1 \text{ mol } CO_2}{44 \text{ g } CO_2} \times \frac{1 \text{ mol } C}{1 \text{ mol } CO_2} \times \frac{12 \text{ g } C}{1 \text{ mol } C} \times 100\% = 60\% \text{ Carbon}$$

$$\frac{14.32 \text{ g H}_2\text{O}}{20 \text{ g sample}} \times \frac{1 \text{ mol H}_2\text{O}}{18 \text{ g H}_2\text{O}} \times \frac{2 \text{ mol H}}{1 \text{ mol H}_2\text{O}} \times \frac{1 \text{ g H}}{1 \text{ mol H}} \times 100\% = 8\% \text{ Carbon}$$

$$60\%C + 8\%H + 32\%O = 100\%$$

$$\begin{aligned} \longrightarrow 100 \text{ g sample} &= 60 \text{ g C} + 8 \text{ g H} + 32 \text{ g O} \\ \text{Ratio: } &\left(\frac{60}{12}\right) \text{ mol} + \left(\frac{8}{1}\right) \text{ mol} + \left(\frac{32}{16}\right) \text{ mol} \\ \longrightarrow 5 : 8 : 2 \text{ ratio} &\longrightarrow \text{Empirical Formula Mass} = 100 \text{ g/mol} \\ &\left(\frac{\text{Molecular Mass}}{\text{Empirical Mass}}\right) = 2 \\ \longrightarrow \text{Molecular Formula} &= 2 \times (\text{Empirical Formula}) \\ \longrightarrow \text{Molecular Ratio: } &10 : 16 : 4 = \text{C}_{10}\text{H}_{16}\text{O}_4 \end{aligned}$$

- **Precipitation Reactions**

In an aqueous solution, precipitation is the process of transforming a dissolved substance into an insoluble solid from a super-saturated solution. The solid formed is called the precipitate.

Aqueous Solution (*aq*) - A solution in which the solvent is water.

**Continue the Notes:**

**Precipitation Reactions**

**Molarity**

Solution = Solvent + Solute  
Solvent - The 'matrix', most often the one in larger amount that integrates the solute physically in it  
Solute - The substance in lesser supply that 'disappears' or dissolves into the solvent

Solvation - The process of dissolving  
Hydration - Dissolving by water

Features of Solution:

- Transparency (despite color)
- Homogeneity

Compare with Mixtures and Suspensions:

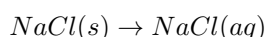
- Mixture - two or more substances that are not chemically combined.
- Suspension - a heterogeneous mixture of a finely distributed solid in a liquid. The solid is not dissolved in the liquid, as is the case with a mixture of salt and water.
  - Milk is a fat suspension and a sugar solution.
  - Milk is an emulsion - a mixture of two or more liquids that are normally immiscible owing to liquid-liquid phase separation.
  - Emulsions are part of a more general class of two-phase systems of matter called colloids.

**Water As a Solvent Dissolving a Salt**

Multiple  $\delta^-$  Oxygen ends surround a Cation. Multiple  $\delta^+$  Hydrogen ends surround an Anion.

Properties of Water:

- Polar Molecule:  $\mu \neq 0 \iff (2) \text{ Hydrogen} = \delta^+, (1) \text{ Oxygen} = \delta^-$



If the speed of Ionic Dissociation is Fast, it is called a Strong Electrolyte. If the speed of Ionic Dissociation is Slow, it is called a Weak Electrolyte.

**Concentration**

- Concentrate - a mixture with relatively more solute in solution
- Dilute - a mixture with relatively less solute in solution

- **Solid in Liquid:** Sugar in Water
- **Gas in Liquid:** Carbonated Water
- **Liquid in Liquid:** Alcoholic water cocktail mixers
- **Solid in Solid:** Alloys are metallurgical homogenous mixtures
- **Molarity Formula & Stock Solutions → Produce Dilutes**

$$\text{constant: } n(\text{moles}) = C \times V$$

$$C_1 V_1 = C_2 V_2$$

- **Mass % (Weight/Weight):**

$$\left( \frac{\text{Mass of Solute}}{\text{Total Mass of Solution}} \right) \times 100\%$$

Also Mass-Volume Percentage (w/v)%

- $\rho(\text{H}_2\text{O}) = 1\text{g/cm}^3$

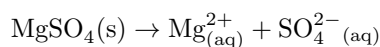
$$1\text{L} = 1\text{kg} = 1,000,000\text{mg}$$

- **Parts-per-trillion (ppt):**

$$\text{PPT} = \left( \frac{\text{Mass of Solute}}{\text{Total Mass of Solution}} \right) \times 10^{12}$$

- **CLC: Dr. G's Wine Question?**

- **Example:**



How many milliliters of 0.5M Magnesium Sulfate Solution must be added to an existing volume of 165ml pure water to achieve a solution that is 0.2M with regards to Sulfate Ion Concentration?

$$C_1 \times V_1 = C_2 \times V_2$$

$$(0.5) \times (V_1) = (0.2) \times (V_1 + 165)$$

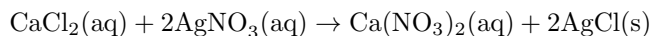
$$0.5 \times V_1 = 0.2V_1 + .033$$

$$0.3V_1 = 0.033$$

$$V_1 = 110\text{ mL}$$

- **Precipitation Reaction:**

- **Complete and Net Ionic Reactions:**



**Express every Ionic Species as Dissociated Components**

⇒ Complete Ionic Equation ⇒ Cancel out those which appear in both Products and Reactants

⇒ Express Net Ionic Equation without Spectator Ions

- **Solubility Rules etc.**

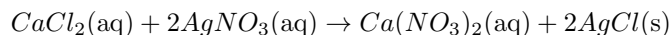
- **Exam:**

- Formal Charge
- Electron Geometry
- Partial Charge

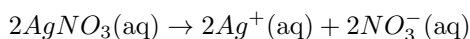
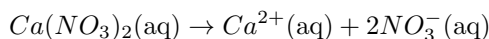
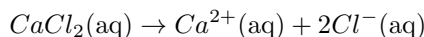
- Ionic Bond & Covalent bond & Metallic Bonds & IMF rankings (I.D. & H & dd & induced D & LDF)
- Boiling Point and Vapor Pressure Comparison
  - B.P.: Honey & Water & Whiskey
  - V.P.: Whiskey & Water & Honey
- Assume non-polar if entirely Carbon-Hydrogen OR completely symmetrical
  - ex.  $\text{CH}_3\text{CH}_2\text{CH}_2\text{CH}_3$  has a very low boiling point
- Chemical Ether properties and IMF/boiling
- Singular Metals take more energy to boil
- Larger LDFs are stronger - more overlap
- Smaller H-bonds are Stronger - more networking
- Electron Protrusions' Repulsion Effect on Molecular Geometry Angles
  - Octahedral's Square Planar  $\text{AX}_4\text{E}_2$  is the same as the Linear  $\text{AX}_2\text{E}_4$  - 90° and 180° degrees
  - $\text{CCl}_3$  is unstable when it is alone it has 120°
  - $\text{NF}_3$  has a lone pair & 109.5°
  - $\text{NH}_4^+ = 109.5$
- Electron Hybridization → Bond Vectors and Sigma-Pi ( $\sigma - \pi$ ) Bonds
- Formal Charge Stabilization is Better than Octet Completion
- Combustion Analysis Steps
- Limiting Reactant Steps

## 1 chem\_102-03.23.2023-notes: Electrodeposition

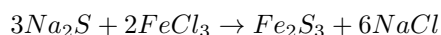
Ex.



Soluble = Dissociate in Solution



Ex.



65% yield, how much 0.200M Na<sub>2</sub>S needed to produce 0.345g Fe<sub>2</sub>S<sub>3</sub>?

$$\text{MM}(\text{Na}_2\text{S}) = 78.045 \text{ g/mol}, \text{MM}(\text{FeCl}_3) = 162.204 \text{ g/mol}$$

$$\text{MM}(\text{Fe}_2\text{S}_3) = 207.885 \text{ g/mol}, \text{MM}(\text{NaCl}) = 58.443 \text{ g/mol}$$

$$\frac{0.345 \text{ g}}{0.65} = 0.53077 \text{ Fe}_2\text{S}_3 = 0.0025532 \text{ moles} \rightarrow 0.0076596 \text{ Na}_2\text{S}$$

$$\frac{n(\text{Na}_2\text{S})}{C[\text{Na}_2\text{S}]} = \left( \frac{0.0076596}{0.200} \right) \times 1000 \approx 38.298 \text{ ml}$$

### Solubility Rules

- Spectator Ions and Ionic Equations

### Separation/Filtering Based on Individual Solubility

#### Acid-Base Reactions

- Strong Acids/Bases
- Weak Acids/Bases

- Neutralization and Equivalent/Reversible Reactions
- Net Ionic Equations

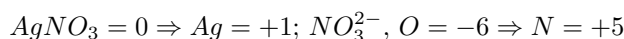
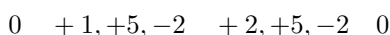
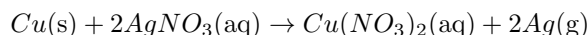
### Redox Reactions and Oxidation Number

- OIL RIG - Oxidation is Loss, Reduction is Gain of Electrons

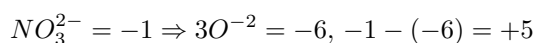
Facilitating Agent:

- Reducing Agent = Species undergoing Oxidation (reduces others, oxidizes itself)
- Oxidizing Agent = Species undergoing Reduction (oxidizes others, reduces itself)

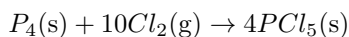
Ex.



OR



Ex.



P is the reducing agent

Can every reaction be classified as a REDOX reaction?

- No
- $AgNO_3 + NaCl \rightarrow AgCl + NaNO_3$
- $+1, -1 \quad +1, -1 \quad +1, -1 \quad +1, -1$

$$\text{Pressure} = \frac{\text{Force}}{\text{Area}}$$

$$\text{Pressure Units } \left( \frac{N}{m^2} \right) = \frac{\text{Newtons (N)}}{\text{Square Meters (m}^2\text{)}} \quad \text{Elephant Pressure} = \frac{14,000 \text{ lbs}}{1000 \text{ in}^2} = 14 \text{ lb/in}^2 \quad \text{Skater Pressure} = \frac{120 \text{ lbs}}{4 \text{ in}^2} = 30 \text{ lb/in}^2$$

### Kinetic Molecular Theory: Work/Pressure Unit Relationships

#### Atmospheric Pressure

#### Devices: Barometer vs Manometer

- Encapsulation vs Open Gases
- Hydrostatic Pressure (P) =  $h \propto P \propto g$

Pressure, Volume, Temperature, Work, Relationships: [https://chem.libretexts.org/Bookshelves/Physical\\_and\\_Theoretical\\_Chemistry\\_Textbook\\_Maps/Supplemental\\_Modules\\_\(Physical\\_and\\_Theoretical\\_Chemistry\)/Physical\\_Properties\\_of\\_Matter/States\\_of\\_Matter/Properties\\_of\\_Gases/Gas\\_Laws/Gas\\_Laws%3A\\_Overview](https://chem.libretexts.org/Bookshelves/Physical_and_Theoretical_Chemistry_Textbook_Maps/Supplemental_Modules_(Physical_and_Theoretical_Chemistry)/Physical_Properties_of_Matter/States_of_Matter/Properties_of_Gases/Gas_Laws/Gas_Laws%3A_Overview)

#### Gas Laws: 1662, 1787, 1811

- Boyle's Law:  $P \propto \frac{1}{V}$ ;  $P_1V_1 = P_2V_2$
- Charles' Law:  $V \propto T$ ;  $\frac{V_1}{T_1} = \frac{V_2}{T_2}$
- Avogadro's Law:  $V \propto n$ ;  $\frac{P_1}{n_1} = \frac{P_2}{n_2}$
- Ideal Gas Law:  $V \propto \frac{nT}{P} \Rightarrow V = \frac{RnT}{P} \Rightarrow PV = nRT$
- $R = \frac{PV}{nT}$
- $R = \text{gas constant} = 8.3145 \text{ Joules} \cdot \text{mol}^{-1} \cdot \text{K}^{-1} \text{ (SI Unit)} = 0.082057 \text{ L} \cdot \text{atm} \cdot \text{K}^{-1} \cdot \text{mol}^{-1}$

#### Ideal vs Real Gas: ???

- A gas that behaves ideally is an "ideal gas"; a hypothetical construct
- Gases we use daily deviate from ideal behavior; a "real gas"
- They have InterMolecular Forces
- Become 'ideal' under high Temperature and low Pressure  $\Rightarrow$  Effect, not force itself, of IMFs is diminished

$$n(\text{moles}) = \frac{m(\text{mass})}{\text{MM}(\text{molar mass})}$$

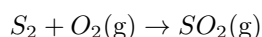
$$\text{density} = \frac{\text{mass}}{\text{Volume}}$$

$$P = \frac{dRT}{M}$$

$$M = \frac{dRT}{P}$$

**Ex.**

Vol of  $\text{SO}_2$  produced at  $343^\circ\text{C}$  and 1.21 atm by 1 kg solid Sulfur?



$$1000 \text{ g S} = 31.187 \text{ mole} = 31.187 \text{ mol SO}_2 (1998 \text{ g SO}_2)$$

$$PV = nRT \Rightarrow V = \frac{nRT}{P} \Rightarrow \frac{(31.187) \cdot 0.08206 \cdot (343 + 273.15)}{1.21} \approx 1303.18536885 \text{ L}$$

**Ex.**

Hydrocarbon gas is 81.7% C by mass.  $T = 25^\circ\text{C}$ ,  $P = 1 \text{ atm}$ , 1.28 g of this gas occupies 0.714 L. Molecular formula?

$$n = \frac{PV}{nRT} \approx \frac{1 \cdot 0.714}{0.08206 \cdot 298} \approx 0.029 \text{ mol}$$

$$\frac{1.28 \text{ g}}{0.029 \text{ mol}} \approx 44 \text{ g/mol}$$

$$\text{MM}(\text{C}_3\text{H}_8) \approx 44 \text{ g/mol}$$

$$\frac{81.7\% \text{ C}}{18.3\% \text{ H}} \Rightarrow 1 : 2.6 \approx 1 : 3$$

$$\frac{100 \cdot 3(12.10)}{3(12.01) + 8(1.008)} \approx 81.7\%$$

**?Dalton's Law of Partial Pressures**

**?Vapor Pressure**

**?Dry gas vs Wet Gas Pressure**

**?water to wine to beer to milk experiment**

**The Kinetic-Molecular Theory KMT - a simple microscopic model that effectively explains the gas laws**

1. Gases are made up of particles with no defined volume but with a defined mass. In other words, their volume is minuscule compared to the distance between themselves and other molecules.
2. Gas particles undergo no intermolecular attractions or repulsions. This assumption implies that the particles possess no potential energy and thus their total energy is simply equal to their kinetic energies.
3. Gas particles are in continuous, linear, random motion, changing direction upon collision with particles or walls.
4. Particle collisions are completely elastic, i.e., no net loss or gain of kinetic energy when particles collide.

5. The average kinetic energy of the gas molecules is proportional to the Kelvin temperature of the gas, regardless of the identity of the gas. Furthermore, this kinetic energy is proportional to the absolute temperature of the gas.

$$KE = \frac{1}{2}mv^2 \quad ; \quad KE = \frac{3}{2}k_bT$$

$$k_b = 1.381 \times 10^{-23} \text{ kg m}^2 \text{ s}^{-2} \text{ K}^{-1}$$

This equation says that the speed of gas particles is related to their absolute temperature. As their temperature increases, their speed increases, and finally, their total energy increases as well. However, it is impossible to define the speed of any one gas particle. As such, the speeds of gases are defined in terms of their root-mean-square speed.

#### KMT (Bhagya Gunasekera)

- Gas molecules/atoms are in continuous linear random motion.
- Gas molecules/atoms are treated as point spaces which have negligible internal unit volume ( $V$ ).
- Collisions between molecules/atoms and with the vessel wall exert pressure ( $P$ ).
- Gas molecules/atoms exert no attractive forces on each other or on the vessel wall.
- Collisions are elastic.

#### Effusion and Diffusion of Gases - Graham's Law

- The mean free path is the average distance a molecule travels between collisions.
- The mean free path increases with decreasing pressure.
- In general, the mean free path for a gaseous molecule will be hundreds of times the diameter of the molecule.

**Process by which molecules disperse in space in response to differences in concentration is called diffusion.**

$$\text{rate of diffusion} = \frac{\text{amount of gas}}{\text{unit of time}}$$

A process involving movement of gaseous species similar to diffusion is effusion; movement of a gas through a small space.

**Graham's law of effusion:** The rate of effusion of a gas is inversely proportional to the square root of the mass of its particles:

$$r \propto \frac{1}{\sqrt{M}}$$

$$\frac{\text{effusion rate A}}{\text{effusion rate B}} = \frac{\sqrt{M_B}}{\sqrt{M_A}}$$

**Velocity is a Vector:** has magnitude and direction.

- **Root-Mean Square Velocity**

$$v_{RMS} = \sqrt{\frac{3RT}{m_{\text{particle}}}}$$

$$R = 8.314 \text{ J K}^{-1}$$

- Mass in Kilograms.
- Gases do not have internal mechanisms for controlling intrinsic velocity; They are driven by environmental temperature.

$$KE_{\text{avg}} = \frac{3}{2}RT$$

#### Molecular Speed Distribution



- Graphs of different substances? Different Temperatures? Trends?

**Ex.** Helium balloon deflates to 2/3 volume in 8 hrs. Time for the same balloon of air to deflate to 1/2 volume?

$$\frac{\Delta n_{\text{He}}/\text{time}}{\Delta n_{\text{air}}/\text{time}} = \frac{\sqrt{28.2}}{\sqrt{4}}$$

$$\frac{(n - \frac{2}{3}n)/8}{(n - \frac{1}{2}n)/t} = \frac{\sqrt{28.2}}{\sqrt{4}}$$

$$\frac{(n \cdot \frac{1}{3})/8}{(n \cdot \frac{1}{3})/8}$$

$$\frac{n}{24} \div \frac{n}{2t} = \frac{\sqrt{28.2}}{\sqrt{4}}$$

$$= \frac{\sqrt{28.2}}{\sqrt{4}}$$

$$PV = nRT$$

$$\frac{PV}{nRT} = 1$$

Compressibility Factor of Ideal Gas = 1 Compressibility Factor of Ideal Gas < 1

### Nature of Chemical Energy

- Energy - the ability to do Work or supply Heat.
- Temperature probe always reports the surrounding temperature, not the internal temperature of the chemical system.

### Calorimetry

- Energy necessary to raise the temperature of 1g of water by 1°C.

### Heat Capacity

- The amount of heat ( $q$ ) a body of matter absorbs or releases when undergoing a 1°C (or 1°K) temperature change.

$$C = \frac{q}{\Delta T}$$

... ???

**Specific Heat Capacity** - the amount of heat ( $q$ ) 1kg of substance absorbs or releases when undergoing a 1°C (or 1°K) temperature change.

$$C = \frac{q}{m \cdot \Delta T}$$

**Units for C:** ???

**Measuring Energy** ( $^{\circ}\Delta$ )

- Specific heat ex.

$$q = m \cdot C \cdot \Delta T$$

$$1.43 \times 10^3 \text{ J} = 217 \text{ g} \cdot C \cdot (39.1^{\circ}\text{C} - 24.5^{\circ}\text{C})$$

$$C = 0.451 \text{ J} \cdot ^{\circ}\text{C}^{-1} \cdot \text{g}^{-1}$$

**Type of Calorimeters - maintain thermal energy on internal contents**

- Coffee Cup - simple and pragmatic.
- Bomb Calorimeter - maintains constant volume.

$$q_{\text{sys}} = -q_{\text{surr}}$$

$$|q_{\text{sys}}| = |q_{\text{surr}}|$$

Ex.

$$q_{\text{sys}} = 248 \text{ g Cu} \cdot 0.385 \text{ J/g}^\circ\text{C} \cdot (T_{\text{f, sys}} - 314^\circ\text{C})$$

$$q_{\text{surr}} = (390 \text{ ml} \cdot 1 \text{ g/ml}) \cdot 4.18 \text{ J/}^\circ\text{C}^{-1} \cdot \text{g}^{-1} \cdot (T_{\text{f, surr}} - 22.6^\circ\text{C})$$

$$q_{\text{sys}} = -q_{\text{surr}}; T_{\text{f, sys}} = T_{\text{f, surr}}$$

$$-1 \cdot 248 \text{ g Cu} \cdot 0.385 \text{ J/g}^\circ\text{C} \cdot (T_{\text{f, sys}} - 314^\circ\text{C}) = 390 \text{ g H}_2\text{O} \cdot 4.18 \text{ J/}^\circ\text{C}^{-1} \cdot \text{g}^{-1} \cdot (T_{\text{f, surr}} - 22.6^\circ\text{C})$$

$$248 \cdot 0.385 \cdot (314 - T_2) = 390 \cdot 4.18 \cdot (T_2 - 22.6)$$

$$29981 - 95.48 \cdot T_2 = 1630 \cdot T_2 - 36843$$

$$T_{2, \text{final}} \approx 38.7^\circ\text{C}$$

Ex.

22.1 g of Al, molar heat capacity = 24.03 J/ C · mol, heated to 82.4°C dropped into calorimeter of water

Water specific heat capacity = 4.18 J/g°C, initially at 22.3°C

Final water temp = 25.8°C

mass of water?

$$n = \frac{-\left(\frac{22.1 \text{ g}}{27 \text{ g/mol Al}} \cdot 24.03 \text{ J/ C} \cdot \text{mol} \cdot (25.8^\circ\text{C} - 82.4^\circ\text{C})\right)}{\left(\frac{4.18 \text{ J/g}^\circ\text{C}}{18 \text{ g/mol H}_2\text{O}} \cdot (25.8^\circ\text{C} - 22.3^\circ\text{C})\right)}$$

$$n \approx 4.22748310169$$

$$n \cdot 18 \text{ g/mol} \approx 76.0946958305$$

## 2 Heat-Work-Volume Relation and Calculation

Enthalpy of Reaction, Heat of Reaction, Specific Heat, Latent Heat, Heat of Fusion, Heat of Vaporization  
Enthalpy of Combustion, Enthalpy of Formation, Hess's Law

$$W = P\Delta V = \frac{N}{m^2} = Nm = J$$

∴ PV work represents Energy

$$0.08206 \text{ L} \cdot \text{atm} \cdot \left(\frac{\text{mol}^{-1}}{\text{K}}\right)$$

$$8.314 \frac{\text{J}}{\text{mol}^{-1} \cdot \text{K}}$$

$$\frac{8.314}{0.08206} = \frac{\text{J}}{\text{L} \cdot \text{atm}} = \frac{W}{P\Delta V}$$

Internal Energy ( $U$ ) is a State Function Paths X and Y represent two different routes to the summit  
Both have the same change in elevation (State Function) Different distances traveled (Path Function)  
(credit: modification of work by Paul Shaffner)

Isolated system vs Open System???

Physics Approach to Internal Energy State Function ???

Difference between Internal Energy and Enthalpy: Internal energy is the total energy in the system  
 $\Delta U = Q + W$  The internal energy is the sum of the heat added to the system and the work done by the system  
Enthalpy is the relationship between the system and the surrounding

$$H = \Delta U + P\Delta V \implies H = \Delta(Q + W) + P\Delta V$$

$\Delta H = q_p$  Enthalpy of Pressure?

For an ideal gas: Internal Energy:  $U = U(T)$  Enthalpy:  $H = U + PV$

$$H = U + RT$$

Specific heats at constant pressure:  $dU = C_v(T)dT$  Specific heats at constant volume:  $dH = C_p(T)dT$   
Specific heat ratio  $k$ :

$$k = \frac{C_p}{C_v} = \frac{U}{H}$$

Work Done by a System:

$$W = \int (F \cdot dx) = P \cdot \int (A \cdot dx) = \int (P \cdot dV)$$

Work done BY a system  $\Rightarrow$  expansion  $\Rightarrow -W$  Work done ON a system  $\Rightarrow$  contraction  $\Rightarrow +W$

Extensive Property vs Intensive Property

Temperature Probe report will ALWAYS be SURROUNDING TEMPERATURE TEMPERATURE  
NEVER refers to the system  $\Delta T_{meter} = \Delta H_{sys}$

Ex. Zn

Enthalpy of Combustion/Formation Enthalpy change when 1mol of substance burns completely

If a process can be written as the sum of several stepwise processes, the enthalpy of change the overall process equals the sum of the enthalpy changes of the steps

Hess's Law ex.

ΩΩΩΩΩΩΩΩΩΩΩΩΩΩΩΩ

Missing: 04/14/2023 Notes

ΩΩΩΩΩΩΩΩΩΩΩΩΩΩΩΩ

### 3 Chemical Equilibria - Understanding Reversibility

Normal: Forward Reaction is 'spontaneous' and reverse reactions is 'non-spontaneous'

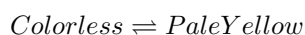
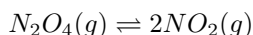
Theoretically, every reaction is reversible Reversibility - Forward and Reverse reactions are both spontaneous

Reversibility Combined with Rate of Reaction Equilibrium - state of system in which the Forward and Reverse reactions occur at the same rate

Importance of Dynamism? In a chemical equilibrium, the forward and reverse reactions do not stop, rather they continue to occur at the same rate, leading to constant concentrations of the reactants and the products.

We can detect a state of equilibrium because the concentrations of reactants and products do not appear to change. However, it is important that we verify that the absence of change is due to equilibrium and not to a reaction rate that is so slow that changes in concentration are difficult to detect.

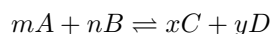
We use a double arrow when writing an equation for a reversible reaction. Such a reaction may or may not be at equilibrium.



When in equilibrium, the solution appears Dark Yellow

REACTION RATE HAS TO BE EQUAL, NOT AMOUNT OF AGENTS

Consider the reversible reaction:



Reactions Quotient:

$$Q_c = \frac{[C]^x \cdot [D]^y}{[A]^m \cdot [B]^n}$$

When  $Q_c = K$ , then the system is in equilibrium

The reaction quotient is equal to the molar concentrations of the products of the chemical equation (multiplied together) over the reactants (also multiplied together), with each concentration raised to the power of the coefficient of that substance in the balanced chemical equation.

The numeric value of  $Q_c$  for a given reaction varies; it depends on the concentrations of products and reactants present at the time when  $Q_c$  is determined. When pure reactants are mixed,  $Q_c$  is initially zero

because there are no products present at that point. As the reaction proceeds, the value of  $Q_c$  increases as the concentrations of the products increase and the concentrations of the reactants simultaneously decrease (Figure 13.6). When the reaction reaches equilibrium, the value of the reaction quotient no longer changes because the concentrations no longer change.

The reaction quotient ( $Q$ ) measures the relative amounts of products and reactants present during a reaction at a particular point in time. The reaction quotient aids in figuring out which direction a reaction is likely to proceed, given either the pressures or the concentrations of the reactants and the products.

When a mixture of reactants and products of a reaction reaches equilibrium at a given temperature, its reaction quotient always has the same value. This value is called the equilibrium constant ( $K$ ) of the reaction at that temperature. As for the reaction quotient, when evaluated in terms of concentrations, it is noted as  $K_c$ .

$$Q_c \text{ at equilibrium} = K_c = \frac{[C]^x \cdot [D]^y}{[A]^m \cdot [B]^n}$$

This equation is a mathematical statement of the Law of Mass Action: When a reaction has attained equilibrium at a given temperature, the reaction quotient for the reaction always has the same value.

Starting with a 1:1 R:P—full:empty reaction, at some point, the rate of transfer will be the same, and both sides will have the same amount. With a 2:1 R:P—full:empty reaction, at some point, the rate of transfer will be the same, yet the end amounts will be different.

Reactant Favored vs Product Favored Equilibrium

$$Q_c < 1 \Rightarrow \text{Reactant Favored}$$

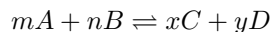
$$Q_c > 1 \Rightarrow \text{Product Favored}$$

$$Q_c < K_c \Rightarrow \text{Forward Reaction until } Q_c = K_c$$

$$Q_c > K_c \Rightarrow \text{Reverse Reaction until } Q_c = K_c$$

### 3.1 Homogeneous Equilibria

Reaction systems where all reactants and products are of the same phase. Versions of Equilibrium Constant;  $K$



$$K_c = \frac{[C]^x \cdot [D]^y}{[A]^m \cdot [B]^n}$$

if all are gases;  $K_p = [P_{C(g)}^x \cdot P_{D(g)}^y] \div [P_{A(g)}^m \cdot P_{B(g)}^n]$

$$P = \frac{n}{V}RT = MRT; \quad P_A = M_A \cdot RT$$

$$K_p = \frac{[P_{C(g)}^x \cdot P_{D(g)}^y]}{[P_{A(g)}^m \cdot P_{B(g)}^n]}$$

$$K_p = ([C] \cdot RT)^x \cdot ([D] \cdot RT)^y \div ([A] \cdot RT)^m \cdot ([B] \cdot RT)^n$$

$$K_p = K_c \cdot (RT)^{(x+y)-(m+n)}$$

$$K_p = K_c \cdot (RT)^{\Delta n}$$

$$Q_p < K_p \Rightarrow \text{Reactant Favored; Forward Reaction until } Q_p = K_p$$

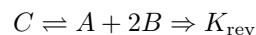
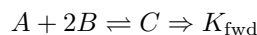
$$Q_p > K_p \Rightarrow \text{Product Favored; Reverse Reaction until } Q_p = K_p$$

### 3.2 Heterogeneous Equilibria

\*\*\*Writing  $K_c$  and  $K_p$  when compounds in a reaction are not in the same phase. Calculation of Equilibrium Constant I-C-E Table Method?

### 3.3 Reaction Synthesis Relation

Ex.



$$K_{\text{fwd}} = \frac{[C]^2}{[A][B]^2}$$

$$K_{\text{rev}} = \frac{[A][B]^2}{[C]^2}$$

$$\therefore K_{\text{fwd}} = \frac{1}{K_{\text{rev}}}$$

$$\therefore K \times K' = 1$$

### 3.4 Shifting Equilibria: Le Châtelier's Principle

When a system is at equilibrium, it experiences a change (in concentration, temperature, pressure, etc.) and it will re-adjust to a new equilibrium state

If a system at equilibrium is subjected to a perturbation or stress (such as a change in concentration), the position of equilibrium changes. Since this stress affects the concentrations of the reactants and the products, the value of  $Q$  will no longer equal the value of  $K$ . To re-establish equilibrium, the system will either shift toward the products (if  $Q < K$ ) or the reactants (if  $Q > K$ ) until  $Q$  returns to the same value as  $K$ .

Ex: Effect of Change in Concentration on Equilibrium ???

Ex: Effect of Change in Pressure on Equilibrium ???

Ex: Effect of Change in Temperature on Equilibrium Cold environments are best to remove heat

$\therefore$  For exothermic reactions, cold environments help remove internal heat and prevent reaction arrest

$\therefore$  For endothermic reactions, hot environments are good to add input reaction heat

???

Catalysts Do Not Affect Equilibrium. A catalyst can speed up the rate of a reaction which may cause a system to reach equilibrium more quickly; a catalyst has no effect on the value of an equilibrium constant nor on equilibrium concentrations. ???

Quadratic Solutions and Extremely Small Values Determining Relative Changes in Concentration Calculations Involving Equilibrium Concentrations

## 4 Exam 3 Review

$$\Delta H_{\text{rxn}} = \sum D_{\text{broken bonds}} - \sum D_{\text{formed bonds}}$$

$$\Delta H_{\text{rxn}} = \sum D_{\text{products}} - \sum D_{\text{reactants}}$$

$$\text{EffusionRate}_A \div \text{EffusionRate}_B = \sqrt{MM_B} \div \sqrt{MM_A}$$

Enthalpy of Formation:  $\text{Mg}_{(s)} + \text{Cl}_2 \rightarrow \text{MgCl}_2(s)$  Activation/Activated Lattice Energy:  $\text{Mg}_{\text{gas}}^{2+} + 2\text{Cl}_{\text{gas}}^- \rightarrow \text{MgCl}_2(s)$

$$\text{Enthalpy of Formation} + -(\Delta H_{\text{rxn}}) = -(\text{Electron\_Affinity} + \text{Lattice\_Energy})$$

- Enthalpy of Formation  $> 0$
- $\Delta H_{\text{rxn}} < 0$
- Lattice\_Energy  $<$  Electron\_Affinity  $< 0$

Open System:

$$\Delta H = \Delta(Q + W) = q \cdot c \cdot \Delta T - (P\Delta V)$$

$q \cdot c \cdot \Delta T$  and  $-(P\Delta V)$  are independent  $\therefore$  We can use  $\Delta H$

Closed System:  $Q$  and  $W$  are interdependent  $\Rightarrow$  We do not know how much the correction for Work has to be to not affect  $Q$  We cannot firmly state the Enthalpy of the considered system Unable to Expand or Contract  $\Rightarrow$  Thermochemically considered a Closed System  $\therefore$  We can only use  $\Delta E$

## 5 Gas Laws

### 5.1 Gas Laws

- ???
- ???
- ???

### 5.2 Linear

- $V \propto T$
- $P \propto T$
- $V \propto n$

### 5.3 NOT INVERSE $\Rightarrow$ Hyperbolic

#### 5.3.1 Boyle's Law

$P \propto 1/V$   $P$ - $V$  Plot is Hyperbolic

#### 5.3.2 PV diagram for thermodynamic conditions?

## 6 Double-Replacement Reactions?

## 7 Solubility Rules

- ???
- ???
- ???

## 8 Net Ionic Equation?

## 9 Strong and Weak Acid/Bases & Salts

- ???
- ???
- ???

## 10 Temperature Probes are ALWAYS measuring the surrounding

## 11 Redox - Reduction/Oxidation

OIL RIG - Oxidation is Loss, Reduction is Gain of Electrons The Oxidation Agent itself is Reduced  $\Rightarrow$  It oxidizes other chemical species The Reduction Agent itself is Oxidized  $\Rightarrow$  It reduces other chemical species

## 12 PT Graphs, Triple-Point and Phase Diagrams, Density Relations

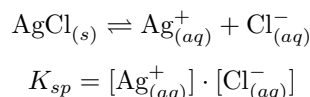
Critical Point, Phase Point, Phase Boundaries}  
 $\Omega$

## 13 Precipitation and Dissolution

### 13.1 Equilibrium of Dissolution

Physical (not Chemical) Equilibrium: Solubility Product =  $K_{sp}$ ?

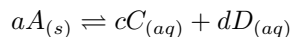
The equilibrium constant for the equilibrium between a slightly soluble ionic solid and a solution of its ions is called the solubility product ( $K_{sp}$ ) of the solid. Recall from the chapter on solutions and colloids that we use an ion's concentration as an approximation of its activity in a dilute solution. For silver chloride, at equilibrium:



When looking at dissolution reactions such as this, the solid is listed as a reactant, whereas the ions are listed as products. The solubility product constant, as with every equilibrium constant expression, is written as the product of the concentrations of each of the ions, raised to the power of their stoichiometric coefficients. Here, the solubility product constant is equal to  $\text{Ag}^{+}$  and  $\text{Cl}^{-}$  when a solution of silver chloride is in equilibrium with undissolved AgCl. There is no denominator representing the reactants in this equilibrium expression since the reactant is a pure solid; therefore  $[\text{AgCl}]$  does not appear in the expression for  $K_{sp}$ .

The solubility product constant,  $K_{sp}$ , is the equilibrium constant for a solid substance dissolving in an aqueous solution. It represents the level at which a solute dissolves in solution. The more soluble a substance is, the higher the  $K_{sp}$  value it has.

The general dissolution reaction (in aqueous solutions):



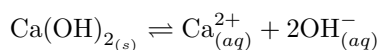
To solve for the  $K_{sp}$  it is necessary to take the molarities or concentrations of the products ( $cC$  and  $dD$ ) and multiply them. If there are coefficients in front of any of the products, it is necessary to raise the product to that coefficient power (and also multiply the concentration by that coefficient). This is shown below:

$$K_{sp} = [C]^c \cdot [D]^d$$

Note that the reactant,  $aA$ , is not included in the  $K_{sp}$  equation. Solids are not included when calculating equilibrium constant expressions because their concentrations do not change the expression; any change in their concentrations is insignificant, and therefore omitted. Hence,  $K_{sp}$  represents the maximum extent that a solid can dissolve in a solution.

### 13.2 Calculations of $K_{sp}$ - Determination of Solubility from $K_{sp}$ via ICE Table

#### 13.2.1 Example



$$\begin{aligned}
\text{Ca}^{2+} &\Rightarrow [ICE] = [0, +s, s] \\
\text{OH}^- &\Rightarrow [ICE] = [0, +2s, 2s] \\
K_{sp} &= [\text{Ca}_{(aq)}^{2+}] \cdot [\text{OH}_{(aq)}^-]^2 \\
1.3 \times 10^{-6} &= (s) \times (2s)^2 \\
1.3 \times ((1.3 \times 10^{-6} = 4s^2) \div 4)^{\frac{1}{3}} \\
s &= (1.3 \times 10^{-6}) \\
s &\approx 6.9 \times 10^{-3} \text{ M}
\end{aligned}$$

$K_{sp}$  = Saturation Value (maximum product of ion concentrations there can be)

### 13.3 Le Châtelier's Principle

Important effects:

1. For highly soluble ionic compounds, the ionic activities must be found instead of the concentrations that are found in slightly soluble solutions.
2. Common Ion Effect: The solubility of the reaction is reduced by the common ion. For a given equilibrium, a reaction with a common ion present has a lower  $K_{sp}$ , and the reaction without the ion has a greater  $K_{sp}$ .
3. Salt Effect (diverse ion effect): Having an opposing effect on the  $K_{sp}$  value compared to the common ion effect, uncommon ions increase the  $K_{sp}$  value. Uncommon ions are ions other than those involved in equilibrium.
4. Ion Pairs: With an ionic pair (a cation and an anion), the  $K_{sp}$  value calculated is less than the experimental value due to ions involved in pairing. To reach the calculated  $K_{sp}$  value, more solute must be added.

#### 13.3.1 Examples

- $2\text{SO}_{2(g)} + \text{O}_{2(g)} \rightarrow 2\text{SO}_{3(l)}$

$$Q = \frac{[\text{SO}_3]^2}{[\text{SO}_2]^2 \cdot [\text{O}_2]} = Q \times 0.640 = K; Q < K \text{ i.e. Forward Reaction}$$

- Lowest Solubility in Water?

$$\begin{aligned}
- \text{Ag}_2\text{S} &= 1.6 \times 10^{-49} \\
&\left(\frac{K}{4}\right)^{\frac{1}{3}} \approx 1 \times 10^{-16} \\
- \text{Bi}_2\text{S}_3 &= 1.0 \times 10^{-72} \\
&\left(\frac{K}{108}\right)^{\frac{1}{6}} \approx 1 \times 10^{-14} \\
- \text{HgS} &= 1.6 \times 10^{-54} \\
&(K)^{\frac{1}{2}} \approx 1 \times 10^{-27} \\
- \text{Mg}(\text{OH})_2 &= 8.9 \times 10^{-12} \\
&\left(\frac{K}{4}\right)^{\frac{1}{3}} \approx 1 \times 10^{-4} \\
- \text{MnS} &= 2.3 \times 10^{-13} \\
&(K)^{\frac{1}{2}} \approx 1 \times 10^{-6}
\end{aligned}$$

$\therefore$  HgS has the Lowest Solubility in Water

The Equilibrium Constant ( $K_p$ ) for a given reaction only has one value per temperature