

On a graduated cylinder with markings on each mL, uncertainty can be measured as 30.5 mL with the .5 being the value estimated while the 30 is certain.

To show the uncertainty, 30.5 ± 0.5 mL . the uncertainty of a graduated cylinder ranges from ±0.1mL to±0.5mL

Error: difference between a measurement and the actual value. Error is usually expressed as a positive value, so there is an absolute value bar.

Absolute Error = |Measurement – Actual Value|

Calculating percent error

A close up of a logo

Description automatically generated

A close up of a logo

Description automatically generated

Systematic error: error that is constant through the set

Sources of systematic error

* Calibration
* Experimental design
* Operating conditions

Random error: error that varies at random from one value to the next in the set

Sources of random error

* Air currents
* Temperature
* Vibrations

Precision: agreement/closeness of repeatable values (can be precise but not accurate)

Accuracy: agreement/closeness of measured values to actual known value (can be accurate overall but not precise)

Significant figures

* When multiplying, we use the fewest of significant figures
* When multiplying, we use the fewest decimals
* Regular rounding rules
* Only round the final answer

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Density is a measure of how atoms pack tightly in a given space

* Density is affected by temperature
* Density is an intensive property (doesn’t change based on amount)
* Mass and volume are extensive properties (changes based on amount)

Temperature: measure of the average kinetic energy of a substance

Conversion factors between temperature scales

* ˚F = 1.8x ˚C + 32
* ˚C = (˚F-32)/1.8
* K = ˚C + 273

Classifying matter

* Classifying by state
  + Solid
    - Fixed volume
    - Fixed shape
    - Can’t flow easily
    - Can’t be easily compressed
  + Liquid
    - Fixed volume
    - Not fixed shape
    - Can flow easily
    - Can’t be easily compressed
  + Gas
    - Not fixed volume
    - Not fixed shape
    - Can flow easily
    - Can be compressed easily
* Classifying by composition
  + Is the composition of matter uniform and constant?
    - Yes, pure substance
    - No, mixture (can be separated into a pure substance)
  + (pure substances) does it consist of one type of atom?
    - Yes, element (like pure oxygen gas)
    - No, compound (like CO [carbon monoxide])
  + (mixtures) are the components in the mixture uniformly distributed?
    - Yes, homogeneous (like salt mixed with water)
    - No, heterogenous (like oil on water)

Physical property: qualities of a substance that can be observed without changer the composition of the substance (color, texture, density, melting point, boiling point, freezing point, particle size)

Chemical property: qualities a substance shows as it changes into or interacts with another substance (gas burning in oxygen)

If the composition of the matter has changed it is a chemical change, if not then it is a physical change

Separating mixtures

* Distillation (takes advantage of differences in boiling points)
* Crystallization (takes advantage of differences in solubility)
* Filtration (takes advantage of difference in particle size)
* Chromatography (differences in solubility)

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Democritus: there is a point where stuff can’t be divided anymore

Dalton’s atomic theory

* Elements consist of small particles called atoms, and these atoms are indivisible
* Atoms of the same element have equal mass. Elements of different elements have different mass
* Different elements usually combine to from a small particle of a molecule
* Hard-sphere model

Thomson’s cathode ray tube experiments

* Cathode ray consists of negatively charged particles called electrons
* Negatively charged particles were part of the atom
* Plum pudding atomic model

Rutherford’s gold-foil experiment

* Discovered nucleus because alpha particles were deflected at weird angles and some were reflected
* Nucleus has protons and neutrons and there was an electron cloud

Bohr’s atomic model

* Electrons move in shells (energy level)

Atomic number (Z): number of protons in an atom/ number of electrons in a neutral atom

Neutron number (N): number of neutrons in an atom

Mass number (A): sum of atomic number and neutron number (A=Z+N)

Isotope: variations of atoms with different mass numbers (because different number of neutrons)

Ion: charged atom (has extra electrons, or less electrons)

Atomic mass is relative to the mass of carbon 12 being 12 amu

Percentage abundance is measured by the proportion of an isotope in a sample from the natural world

Weighted average is the sum of percentage abundance \* mass of isotope

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c = v λ

Where c is the speed of light (3.00\*108 m/s), v is frequency, and λ is wavelength

Wavelength is measured from crest to crest (top of a wave) or trough to trough (bottom of a wave)

Blackbody radiation:

* Things radiate heat at different colors
* Max Planck explained it as heated objects emitting packets of light called quanta

Energy in a quanta

Equanta = h vradiation

Where Equanta is energy in a quanta, h is Planck’s constant (6.63\*10-34 J\*s), and vradiation is frequency of radiation

Using c = v λ, the equation can be rewritten as E = (h\*c)/λ

Einstein explained the photoelectric effect (some metals eject electrons when light is shined on it) by having light be a stream of particles called photons (light has both wave-like and particle-like properties)

Energy of an ejected electron

KEElectron = hv - hvo

Where h is Planck’s constant, v is frequency and vo is threshold frequency (minimum energy required for the photoelectric effect to work)

KEElectron = ½ mv2

Where m is mass, and v is velocity

White light generates a continuous spectrum (not explained)

Light generated from gasses (like hydrogen) have a line emission spectrum because the electrons absorb a specific amount of energy to jump up one (or more) energy levels then emit energy as they fall back down to their regular energy level (this process is called a quantum jump)

Bohr’s atomic model

* Atoms consist of shells or orbits around them
* Orbits are at fixed distances from the nucleus
* Orbits are at a certain energy level from the nucleus
* Electrons move around the nucleus in orbits
* Energy of an electron is connected to the energy of the orbit
* Electrons emit energy in packets or bundles as they jump from one energy level to another

Electrons that are closer to the nucleus are at a lower energy and the atom is more stable

Electrons that are further from the nucleus have more energy and the atom is less stable

Change in energy as electrons travel from Ei (initial energy) to Ef (final energy)

ΔΕ = Ef – Ei = 2.179\*10-18 J (1/n2i – 1/n2f)

Where ΔΕ is change in energy, Ei is initial energy, Ef is final energy, ni is initial principal quantum level (energy level/ shell), and nf is the final principle quantum level

Excited state (of an atom): when its electrons absorb energy and jump from a lower-energy shell to a higher-energy shell

Ground state (of an atom): when all its electrons occupy their lowest possible energy level

Bohr’s atomic model couldn’t explain the spectra of atoms larger than hydrogen

De Broglie suggested that electrons also have wave-like properties

Heisenberg said that it is difficult to find the exact location and momentum of an electron in an atom, so everything had to be a probability

Shrödinger combined all the ideas and made a complex equation

Orbitals

* Where there is a high chance of finding an electron
* s, p, d, f
* How to describe
  + Principal quantum number (n) determines the energy of an electron
    - What energy level (n = 1, 2, 3, …)
  + Azimuthal quantum number (l) describes the shape of the orbital
    - l = n-1
  + Magnetic quantum number (ml) describes the orientation of the orbital
    - ml = (-l,0,+l)
  + Spin quantum number (ms) describes the spin states of an electron
    - ms = +1/2 or -1/2

|  |  |  |  |
| --- | --- | --- | --- |
| Principal energy level (n) | Angular momentum (l) | Magnetic (ml) | Electron spin (ms) |
| 1 (s orbital) | 0 | 0 | +1/2 or -1/2 |
| 2 (p orbital) | 0 | 0 | +1/2 or -1/2 |
| 1 | -1, 0, +1 | +1/2 or -1/2 |
| 3 (d orbital) | 0 | 0 | +1/2 or -1/2 |
| 1 | -1, 0, +1 | +1/2 or -1/2 |
| 2 | -2, -1, 0, +1, +2 | +1/2 or -1/2 |
| 4 (f orbital) | 0 | 0 | +1/2 or -1/2 |
| 1 | -1, 0, +1 | +1/2 or -1/2 |
| 2 | -2, -1, 0, +1, +2 | +1/2 or -1/2 |
| 3 | -3, -2, -1, 0, +1, +2, +3 | +1/2 or -1/2 |

Rules for filling atomic orbitals

* Pauli exclusion principle
  + Each orbital can only have 2 electrons and they have to have different spins
* Hund’s rule
  + When filling orbitals, you have to fill each orbital with one electron before going back to fill the second electron

Parts of the electron configuration (for hydrogen)

1s1 1 = energy level, s = type of orbital, 1 = number of electrons

Noble gas notation (for lithium)

[He] 2s1

Valence electron: electron in the highest energy level, they are used to bond with other atoms

Diamagnetic: electrons are paired

Paramagnetic: has unpaired electrons all

Isoelectronic: has the same structure of valence electrons (electron configuration)

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Atomic radius: one-half of the internuclear distance in the Cl2 molecule

Down and left increase in atomic radius

z\* = z – δ

where z\* is effective nuclear charge, z is the nuclear charge (# of protons) and δ is the shielding effect by the inner electrons

Cations are always smaller than neutral atoms

Anions are always larger than neutral atoms

Ionization Energy: how much energy it takes to remove an electron

Down and left reduces ionization energy

Electron affinity: how willing an atom is to accept an electron

If electron affinity is negative, it is exothermic

If electron affinity is positive, it is endothermic

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Chemical bond

* Intermolecular forces (attractive forces between molecules)
  + Dispersion forces
  + Dipole-dipole forces
  + Hydrogen bonding
* Intramolecular forces (attractive forces within molecules or formula units)
  + Ionic bond (bonds between metals and non-metals) (like NaCl)
    - Metal loses electron
    - Non-metal gains electron
  + Covalent bonds (bonds between non-metals) (Like water)
    - Share electrons
* Chemical bonds happen to lower their energy to become more stable
* When they bond, they make new molecules with unique physical and chemical properties

Monoatomic ion: ions that are of a single atom

Polyatomic ion: charged molecules

Ionic compounds

* Low ionization energy atom with a high electron affinity atom
  + F = (k Q1 Q2) / d2 (attraction between positive and negative ions increases with charge and decreases with distance)
* Naming
  + Name of cation first name of anion+ide (sodium chloride)
* Properties
  + High melting points
  + High boiling points
  + Good conductors of electricity when molten or dissolved in water
  + Poor conductors of heat and electricity when solid
  + Usually soluble in water

Covalent bonds

* Both atoms want more electrons, but they aren’t strong enough to pull the electron off the other atom, so they end up sharing the electron
* Naming
  + Prefix+(name of 1st element) Prefix+(base name of 2nd element)+ide (P2O5 diphosphorus pentoxide)
* Bond polarity and bond types
  + Nonpolar: electrons are shared equally, electronegativity difference of 0.0 – 0.4
  + Polar: electrons are shared unequally, electronegativity difference of 0.4 – 2.0

Bond properties

* Bond length
  + The distance between the nuclei of two boned atoms
  + Bond length increases as period increases (same column but further down)
* Bond energy
  + Energy change when the bond between two atoms are broken and the atoms are completely separated
  + As the number of bonds or electron density between two atoms increases, the bond gets shorter and stronger (bond strength and bond length are negatively correlated)

Exceptions to the octet rule

* They have less than eight electrons (BeF2)
* They have more than eight electrons (SF6) (usually applies when they use the d orbital to bond)
* They have an odd number of electrons (NO the nitrogen has a lone electron)

To break bonds, the atoms involved must absorb energy

To make a bond, the atoms involved must release energy

Because Lewis structures can’t explain magnetic properties, Molecular Orbital theory is a thing

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VESPER

* Using VSEPR (AXE)
  + A: central atom
  + X: atoms bonded to A
  + E: number of lone pairs on the central atom
* Geometrics of bonds
  + Electron pair geometry: number and arrangement of electron pairs around the central atom (name should be the base shape (trigonal planer, tetrahedral, etc.))
  + Molecular geometry: arrangement of the atoms in space
  + Impact of lone pairs
    - Pushes on everything else because it wants to be as far away as possible
    - ^ causes the bond angles between the other bonds to be smaller (tetrahedral has 109.5˚ while water has a bond angle of 104.5 because the lone pairs are pushing those bonds closer together)
* Electrons and lone pairs always want to repel each other so they want to be as far away as possible
  + Relative strength of repulsion: lone pair-lone pair > lone pair-bond > bond-bond
* Drawing the molecule in 3D on paper
  + Straight line: bond is in the plain of the paper
  + Hatched wedge: bond is going into the paper
  + Solid wedge: bond is going out of the paper

Valence bond theory

* Shared electrons are in the overlap region of the orbitals of multiple atoms
* Orbital hybridization: hybrid orbitals bond better
  + One s atomic orbital + one p atomic orbital = two sp orbitals
  + One s atomic orbital + two p atomic orbitals = three sp2 hybrid orbitals
  + One s atomic orbital + three p atomic orbitals + one d atomic orbital = 5 expanded octet orbitals

Types of bonds when orbitals overlap

* Sigma (σ) bond
  + Can from hybridized and unhybridized orbitals
  + Sigma bonds usually form when atomic orbitals overlap along the bond axis
  + Single bonds are usually sigma bonds
  + Part of every bond type
* Pi (π) bond
  + Formed from unhybridized orbitals
  + Pi bonds are usually formed when unhybridized p orbitals overlap side-ways or above and below the bond axis

Molecular polarity

* Polar
  + Permanent dipole: one end is partially positive while the other end is partially negative
  + Dipole moment: an arrow with a plus at one end and an arrowhead at the other, with the + pointing towards the more positive atom and the arrowhead pointing towards the more negative atom +-----> (looks like that)
  + They align when in an electric field
* Non-polar
  + Dipole moments cancel (net dipole movement of zero)

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Intermolecular forces impact phases of matter (solid, liquid, gas)

* Types
  + Dispersion forces: present in all molecules
    - Increases with number of electrons and molar mass
  + Dipole-dipole forces: attraction between polar molecules
    - Higher boiling point relative to dispersion forces
  + Hydrogen bonding: hydrogen being attracted to lone pairs on N, O, F (can be between molecules or within molecules)
    - Greatly increases boiling point

Non-permanent dipoles: electrons randomly happen to be more concentrated on one side of a molecule than the other

Water temperature stays the same for periods where heat is added because heat is used to break hydrogen bonds between water molecules

Phase diagram: a diagram that shows the phases a substance will be at under specific pressure and temperatures. (at the critical point, substances act like both a liquid and gas)

Chemical reaction

* Properties of a chemical reaction
  + Mass is conserved
  + Atoms rearrange to form new chemicals
  + New properties are seen (compared to reactants)
* Signs of a chemical reaction
  + Energy is transferred to/from the surrounding (in the form of heat)
  + Energy is released to the surrounding in the form of light
  + Color change
  + Composition change in the reaction of the mixture (solid, liquid or gas appear)
* Types of reactions
  + Combination (A+B 🡪 AB)
  + Decomposition (AB 🡪 A+B)
  + Single replacement (A+BC 🡪 AC+B)
  + Double replacement (AB+CD 🡪 AD+BC)
    - Factors that drive the reaction
      * Formation of an insoluble ionic compound (precipitation)
      * Formation of a molecular compound that remains in solution
      * Formation of a gaseous compound that escapes from the solution
* Described by a chemical equation (reactants 🡪 products)

Mass-mass stoichiometry

1. Divide given mass by molar mass (convert into g if necessary)
2. Convert ^ to the desired unknown substance using mole ratios (from equation)
3. Multiply ^ by molar mass of unknown substance

Determining limiting reactant

1. Divide reactant masses by their respective molar masses (convert into g if necessary)
2. Divide the moles of each reactant by their respective coefficients
3. The one with the smallest ^ is the limiting reactant

Theoretical yield: amount that can be made if all of the limiting reactant is used

Percent yield = actual yield/theoretical yield \* 100

Getting empirical formula from percent composition

1. Divide % by molar mass (assumes that we have 100g of sample)
2. Divide values from ^ by the smallest number
3. Multiply by a whole number if necessary, to make all the values a whole number

Solute: the chemical that dissolves in the solvent (usually lesser amount)

Solvent: the chemical that dissolves the solute (usually larger amount)

If solute and solvent amounts are equal, then there is no clear solute or solvent in the solution

Major factors that determine if a solute will dissolve in a solvent (entropy, intermolecular forces, energy)

* Solutes and solvents must be able to separate from themselves
* Solvents and solutes must be able to interact with each other
* Both steps above must be favorable

Solubility: maximum amount of solute that can dissolve in 100g of solvent

* Saturated: max amount that can be dissolved
* Unsaturated: less the max amount that can be dissolved
* Supersaturated: more than max that can be dissolved
* Factors that impact solubility
  + Temperature
    - For most solids, increases with temperature
    - For gasses, decreases with temperature
  + Pressure
    - For solids, no change
    - For gasses, increases with pressure

When using a volumetric flask, make sure solute is dissolved before filling to the line

Concentration: amount of solute dissolved in an amount of solvent or solution

Mass percent: mass of solute (g)/ mass of solution (g) \* 100

Volume percent: volume of solute (mL)/ volume of solution (mL) \* 100

Molarity: moles of solute/volume of solution (L)

Properties of acids and bases

* Acids
  + Tastes sour
  + Turns blue litmus paper red
  + Reacts with metals to produce hydrogen gas
  + Reacts with carbonates to produce carbon dioxide
  + Neutralizes bases in an acid-base reaction to produce salt and water
  + Contains a H+ ion (strong acids dissociate completely)
* Bases
  + Tastes bitter
  + Produces aqueous solutions that feel slippery
  + Turns red litmus paper blue
  + Neutralizes acids in an acid-base reaction to produce salt and water
  + Contains an OH- ion (strong bases dissociate completely)
* When combined they neutralize because the ions bond and create water and a salt

Calculating pH pH = -log [H3O+] negative log of the [concentration]

Calculating pOH pOH = -log [OH-]

Bronsted-Lowry definition of acid

* Acids donate a hydrogen ion and bases accept it
* Every acid has a conjugate base and vice versa (Example: HCl (acid), Cl is the conjugate base)

Titration: base is added to an acid by drops until the equivalence point is reached

Equivalence point: volume of base that completely reacts with an acid

Indicator: a dye that changes color when an acid is neutralized

End point: the point past the neutral point at which the indicator changes color

Titration error: the difference between the end point and the equivalence point

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Oxidation: loss electrons (reducing agent)

Reduction: gain elections (oxidizing agent)

Energy: ability to do work

Energy (joule, J)

* Kinetic energy (can be converted to potential energy)
  + Thermal energy
    - Phase changes (movement of the molecules)
* Potential energy (can be converted to kinetic energy)
  + Chemical potential energy
    - Intermolecular forces (Vander Waals forces, dipole-dipole forces)
    - Intramolecular forces (bonds)
* Defined as the amount of kinetic energy a 2kg object has when moving at 1m/s
* Other units
  + 1 calorie (cal) = 4.184 J
  + 1 Calorie (Cal/nutritional calorie) or kilocalorie (kcal) = 1000cal = 4184 J
  + 1 kilowatt-hour (kWh) = 3.60E6 J

System and surroundings

* System: what we are looking at (when burning wood, wood is the system)
* Surrounding: not the system (when burning wood, the air can be the system)
* Types of systems
  + Open: energy and matter can exchange with the surrounding (open fire)
  + Closed: only energy can exchange with the surrounding (candle being burnt out in an air-tight jar)
  + Isolated: neither energy nor matter can exchange with the surrounding (^ when the surrounding is a different jar on the other side of the planet)

Exothermic reaction: system releases energy to the surrounding

Endothermic reaction: system absorbs energy from the surrounding

Thermodynamics: the study of macroscopic properties and how they react to one another

* Pressure, temperature, volume, composition
* Laws
  + 1st Law: Energy can’t be created or destroyed, only converted from one form or another
    - Internal energy = the sum of kinetic energy and potential energy
    - ∆E = Efinal - Einitial
    - ∆Esystem + ∆Esurrounding = 0

State variables/state functions: a system only dependent on its state (what is its pressure, volume, temperature, energy)

Transferring energy

* Work (w): transfer of energy from one mechanical system to another
* Heat (q): transfer of energy as a result of temperature difference
* ∆E = q + w
* If heat/work is transferred into the system, then q/w > 0 (endothermic reactions)
* If heat/work is transferred out of the system, then q/w < 0 (exothermic reactions)

Calculating change in internal energy (∆E)

* ∆E = qp + w
* q can be determined by experimental measurements of temperature
* We need to determine whether w is under constant volume or constant pressure
  + Constant volume
    - Under constant volume, w = -P∆V
    - Because there is constant volume ∆V = 0
    - Therefore, w = 0 and ∆E = qv (with the v telling us that volume was constant)
  + Constant pressure
    - w = -P∆V
    - because w isn’t 0, ∆E = qp + P∆V

Enthalpy (∆H)

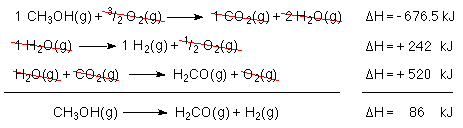
* Basically, temperature change with constant pressure (in terms of the system)
* qp = ∆H = H (products) – H (reactants)
* Endothermic
  + Same # of bonds but weaker
  + Less bonds but same strength
  + ∆H > 0 means it is an endothermic reaction and the system absorbs energy
  + Breaking bonds requires energy (endothermic)
* Exothermic
  + Same # of bonds but stronger
  + More bonds but same strength
  + ∆H < 0 means it is an exothermic reaction and the system releases energy
  + Making bonds releases energy (exothermic)
* Bond Enthalpy
  + Enthalpy in terms of molecules
  + Usually positive because breaking bonds generally takes energy
  + Single bonds are weaker than double bonds which are weaker than triple bonds
* Enthalpy of Reaction
  + ∆Ho = ∑[(moles of bonds)\*D(bonds broken)]- [(moles of bonds)\*D(bonds formed)]
  + Where D is the bond enthalpy
  + Draw Lewis structures to figure out bond structures
  + Example: CH4 + 2 O2 🡪 CO2 + 2H2O
  + ∆Ho = (1\*415(4) + 2\*498(1))-(1\*799(2)+2\*467(2)) = 2656-3466 = -810 kJ
* Measuring experimentally
  + Calorimeter is used to measure heat transfer at either constant pressure or constant volume
  + Because the surrounding is measured, qsystem = -qsurrounding
  + When measuring under constant volume
    - The reaction is done in a bomb calorimeter, which burns the compound in oxygen
    - Because it is done in a bomb calorimeter, qreaction + qbomb + qwater = 0
      * Energy can only be transferred to the calorimeter and water
      * qreaction = -(qbomb + qwater)

Specific Heat

* q = m \* Cs \* ∆T
  + Cs = SH
* Specific heat is the amount of energy needed to raise the temperature of 1g of something by 1˚C
* The unit for specific heat is usually
  + Joules per gram degree Celsius (J/g˚C)
  + Calories per gram degree Celsius (cal/g˚C)

Hess’s law

* Used to find the enthalpy of reactions that can’t be accurately measured using calorimetry
* If the sum of two or more equations can the equation we want, then the ∆H for the equation we want is the sum of the ∆H of the other reactions



Standard molar enthalpy of formation (∆Hf˚)

* ∆H of the formation of one mole of a compound in its standard state (its state at stp)
* ∆Hf˚ of an element at its standard state is 0
* ∆H˚ = ∑[(moles of product) \* ∆Hf˚(product)] - ∑[(moles of reactant) \* ∆Hf˚(reactant)]

Kinetic-Molecular theory of gasses

* Molecules of a gas are much smaller than the distance between them (compressible)
  + - Molecules of gas move randomly at various speeds and in all directions (fills container)
    - Gas molecules do not attract each other
    - Collisions between gas molecules are elastic
    - The average kinetic energy of gas molecules is proportional to the absolute temperature

Pressure

* Force over area
* Atmospheric pressure
  + Pressure exerted by air molecules in the atmosphere (measured by a barometer)
  + 1 atm = 760mm Hg = 760 torr = 29.9in Hg
* Studying gas behavior
  + Pressure (P), Temperature (T [how fast something is moving]), Volume (V), Amount (n)
  + PV=nRT where R is the universal gas constant (R=0.082 (L\*atm)/(mol\*K) is solved by plugging stp into the equation and solving for R)
    - 1 mol of any gas occupies a volume of 22.4L at stp
    - To get K from ˚C, add 273 to ˚C
    - The equation can be converted into P\*mm = D\*R\*T
  + Boltzmann distribution: generally, heavier molecules move slower than lighter ones, but at a given temperature the average kinetic energy of the molecules is the same.
  + Effusion: movement of molecules through a small hole into an empty container
    - Rate of effusion is inversely proportional to its molar mass
      * 1/sqrt(molar mass)

Avogadro’s Law

* At the same temperature and pressure, the ratio of volume to moles of gas is constant
  + If the ratio is 1L to 1 mol then that ration is true for every gas in that situation

Dalton’s Law of Partial Pressure

* Gasses exert pressure called partial pressure
* Total pressure is the sum of partial pressures (PT = PA + PB + …)
* To solve for partial pressure, use PV=NRT

Mole Fraction: moles of a gas divided by the total moles in a gas mixture

* The sum of the mole fractions of a gas mixture must equal 1
* It is unitless
* Partial pressure of gas A = Mole fraction of gas A \* Total pressure

Volume of a gas produced by a reaction can be determined by measuring the amount of water it displaces

* Partial pressure of water must be accounted for and subtracted from the total to get the correct pressure
* The remaining numbers can be used in PV = NRT

Manometer

* If closed ended vacuum: pressure of gas = height of mercury
* If open ended and higher atmospheric pressure: pressure of gas = pressure of atmosphere – height of mercury
* If open ended and lower atmospheric pressure: pressure of gas = pressure of atmosphere + height of mercury

Van Der Waal’s Equation (P+(n^2\*a)/V^2)(V-nb) = nRT

* Where a is the correction for attractive forces and b is the correction for volume of molecules
* This equation is to correct for gasses under high pressure or low temperature where volume and attractive forces become significant.