CHM1025 Exam 3 Review

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Information

- **Lead: Chandler Lentovich, Support: Lam Le**
- **Drop-in tutoring: https://tinyurl.com/ufdropin**
- **Please have a periodic table and calculator out!**
- **There are multiple professors for this class, so you may see material that won't be on your exam during this review**

↑ ↑ ↑ ↑ _

d

Electron Filling Rules

- Aufbau principle
  - Electrons fill orbitals starting with the lowest-energy orbitals
  - [redacted]
- Pauli exclusion principle
  - A maximum of two electrons can occupy each orbital, and they must have opposite spins
- Hund's rule
  - Electrons are distributed into orbitals of identical energy (same sublevel) in such a way as to give the maximum number of unpaired electrons
    - Because negatively charged electrons repel each other

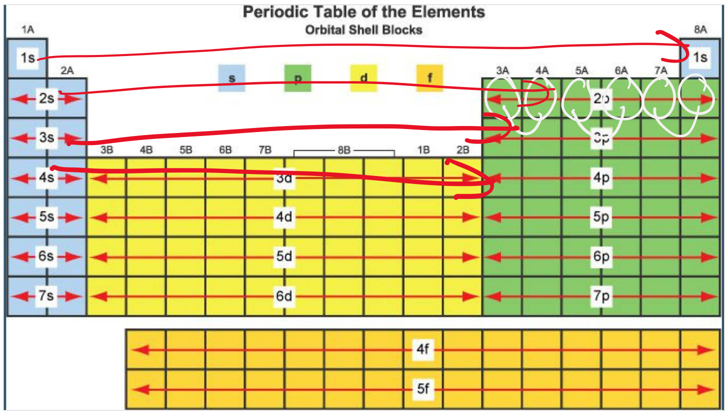
Electron Configuration

- Shows how we fill electrons into their orbitals
- Based on the principles from the last slide $1s^2 2s^2 2p^6 3s^1$
- Example: Na

↑↓ 1s ↑↓ 2s ↑↓ ↑↓ ↑↓ 2p ↑ 3s

- You can also use shorthand, starting with the most recent noble gas
- Example: Br

[Ar] ↑↓ 4s ↑↓ ↑↓ ↑↓ ↑↓ ↑↓ 3d ↑↓ ↑↓ ↑ 4p $[Ar]4s^2 3d^{10} 4p^5$



Periodic Table of the Elements

Orbital Shell Blocks

Special Cases

- It can be more stable to have a half-full s orbital and a completely full d orbital than a full s orbital and almost-full d orbital ← more stable @ higher energy
- Cr down and Cu down follow this rule
- Example: Mo [Kr] ↑ 5s ↑ ↑ ↑ ↑ ↑ 4d
- Ions
  - Add or take away as many electrons as necessary
  - Example: S2- $1s^2 2s^2 2p^6 3s^2 3p^6$

↑↓ 1s ↑↓ 2s ↑↓ ↑↓ ↑↓ 2p ↑↓ 3s ↑↓ ↑↓ ↑↓ 3p

Periodic Properties

- Ionization energy
  - A measure of the energy required to remove a valence electron from an atom to form an ion
    - Increases left to right, decreases up to down
- Atomic size/radius
  - Decreases left to right, increases up to down
  - Ions: become larger with electrons added and smaller with electrons taken away
- Electronegativity
  - A measure of an atom's ability to attract the shared electrons of a covalent bond to itself
    - Increases left to right, decreases up to down

Ionic Bonds

- Ionic= metal + nonmetal
  - Transfer of an electron
- Typically have a high melting and boiling point due to strong electrostatic interactions

Covalent Bonds

- Covalent= Occurs between nonmetals in molecular elements, molecular compounds, and polyatomic ions
  - Electrons shared between atoms in pairs
  - Covalent bonds can be polar or nonpolar
- Polar: significant difference in electronegativity
  - Electrons will be closer to the more electronegative atom (dipole moment)
- Nonpolar: negligent difference in electronegativity (typically, less than 0.4)
  - A compound that has polar bonds but a line of symmetry overall will be nonpolar (ex. CO2, CCl4)

Polarity

- Polar: significant difference in electronegativity (0.5-1.8)
  - Electrons will be closer to the more electronegative atom (dipole moment)
- Nonpolar: negligent difference in electronegativity (0-0.4)
- A compound that has polar bonds but a line of symmetry overall will be nonpolar (ex. CO2, CCl4)
- Dipole arrows
  - Go from positive to negative, "plus" on positive end
  - Example: CH3COOH

Formal Charge

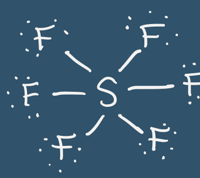
- Formal charge= valence electrons- nonbonding valence electrons- (bonding electrons/2)
- The optimal electron arrangement will have a formal charge of zero
- One way that helps me check:
  - Count each lone pair electron
  - Count one electron from each bond connected to the atom
  - Subtract this amount from the number of valence electrons for that atom
- Example: what is the formal charge on one of the O atoms within CO2?

Lewis Structures

- Step 1: Sum the valence electrons
- Step 2: Determine the center atom (usually the least electronegative atom or carbon if carbon is present)
- Step 3: Draw a "skeleton" using single bonds
- Step 4: Add remaining electrons to complete the octet of each outer atom and then to the central atom if necessary and if there are electrons available
- Step 5: If needed to satisfy the octet rule, shift unshared electron pairs to bonding positions to form double or triple bonds
- Step 6: Draw resonance structures if necessary

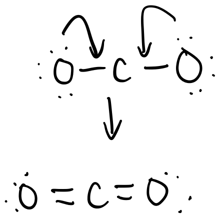
Lewis Structures: SF6

valence
S: 6
F: 7 x 6 = 42
total = 48



bonds = 12 e⁻s
48-12 = 36
36/6 = 6
↑
each F gets 6
3 lone pairs

CO_2
valence
C: 4
O: 6 x 2 = 12
total: 16



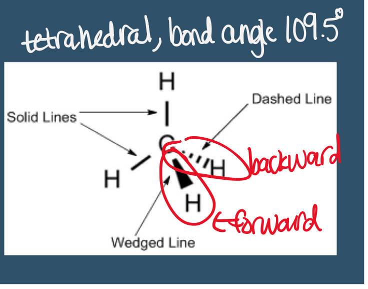
bonds: 4e⁻s
16-4 = 12
2 atoms: 12/2 = 6
3 lone pairs

Wedges and Dashes

- Wedges: forward atom
- Dashes: backwards atom
- Not all atoms get to be in plane!

↑
would all be 90°

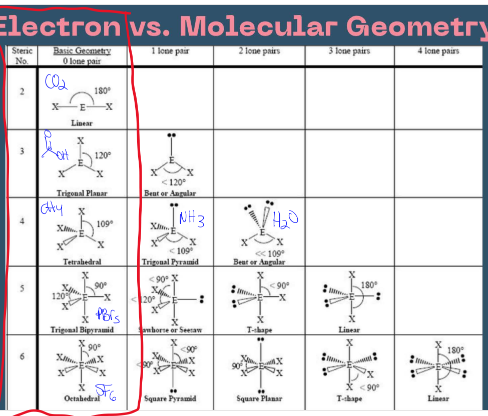
tetrahedral, bond angle 109.5°



Electron vs. Molecular Geometry

areas of electron density

Shape



Steric No.	Basic Geometry 0 lone pair	1 lone pair	2 lone pairs	3 lone pairs	4 lone pairs
2	CO_2 180° Linear				
3	120° Trigonal Planar	< 120° Bent or Angular			
4	CH_4 109° Tetrahedral	NH_3 < 109° Trigonal Pyramid	H_2O << 109° Bent or Angular		
5	120°, 90° PBr_5 Trigonal Bipyramid	< 90°, < 120° Sawhorse or Seesaw	< 90° T-shape	180° Linear	
6	90° SF_6 Octahedral	< 90°, < 90° Square Pyramid	90° Square Planar	< 90° T-shape	180° Linear

Gas Laws

- Boyle's law: as pressure goes up, volume goes down
  - P1V1 = P2V2
- Charles's Law: if temperature (K) increases, volume increases (directly proportional)
  - V1/T1=V2/T2
- Combined gas law
-

$$\frac{P_1V_1}{T_1}=\frac{P_2V_2}{T_2}$$

Ideal Gas Law

- PV=nRT
- Note: 1 mol of an ideal gas at STP occupies 22.4 L
  - STP= 0°C and 1 atm
- Example: The volume of a propane cylinder used in Joel's gas grill is 0.960 L. When the cylinder is "empty," it contains propane gas molecules at atmospheric pressure and temperature. How many moles of propane gas remain in a cylinder when it is empty if the surrounding atmospheric conditions are 25.0°C and 745 torr?

0.0384 mol

V = 0.960 L
P = ~~745 torr~~ 0.980 atm
T = ~~25.0°C~~ 298.2 K
n = ?
$R = 0.0821 \frac{L \cdot atm}{mol \cdot K}$

converting Pressure

$745\,torr \cdot \frac{1\,atm}{760\,torr} = 0.980\,atm$

converting Temperature

25°C + 273.15 = 298.2 k

Ideal Gas Calculation

$$\frac{PV}{RT}=\frac{n\cancel{RT}}{\cancel{RT}} \rightarrow \frac{PV}{RT} = n \longrightarrow \frac{(0.980\,\cancel{atm})(0.960\,\cancel{L})}{\left(0.0821 \frac{\cancel{L}\cdot\cancel{atm}}{mol\cdot\cancel{K}}\right)(298.2\,\cancel{K})} = \boxed{0.0384\,mol}$$

Dalton's Law of Partial Pressures

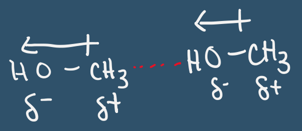
- Gases in a mixture behave independently and exert the same pressure they would exert if they were in a container alone
- The total pressure exerted by a mixture of gases is the sum of the partial pressures of the component gases

London Dispersion Forces

- Temporary dipole that takes place due to an unequal sharing of electrons
- The intermolecular force that is "always there"
- Larger atoms = stronger london dispersion forces

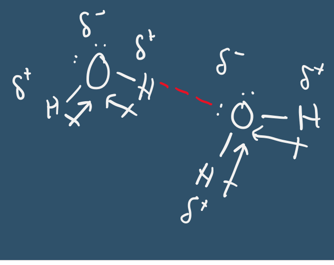
Dipole-Dipole Interactions

- Interactions between two PERMANENT dipoles
- Positive of one molecule attracted to negative of another molecule



Hydrogen Bonding

- H of one molecule interacts with N, O, or F of another molecule ← electronegative
- Think of H as "positive" and N, O, and F as "negative"



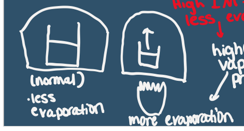
Colligative Properties

- Only vary with concentration of solute particles
- Atoms/molecules with more intermolecular forces acting on them will be more resistant to move to a more disordered state
  - Solid->Liquid->Gas ← molecules further apart
    - Vapor pressure lowering
    - Melting/boiling point elevation
  - Freezing point depression
    - Ions hinder molecules from forming ice crystals

molecules close together

weaker intermolecular forces will move to a more disordered state easier

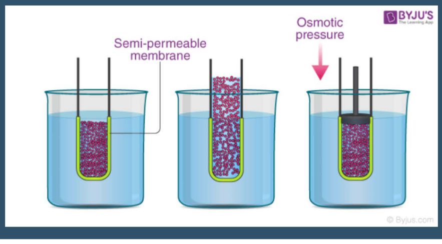
High IM forces = less evaporation



higher vapor pressure ← more evaporation took place ← molecules were overcome by heat, and intermolecular forces overpowered

Osmotic Pressure

- Osmotic pressure: the pressure that would have to be applied to a pure solvent to prevent it from passing into a given solution by osmosis, often used to express the concentration of the solution



Solutions Basics

- Solute is what goes into the solvent
  - U comes before V in the alphabet
- "Like dissolves like"
- Levels of saturation
  - Unsaturated= less than the maximum allowable amount of solute dissolved
  - Saturated= maximum allowable amount of solute dissolved
  - Supersaturated= more than the maximum allowable amount of solute dissolved
    - How to make: make a saturated solution at a lower temperature, then heat it up to increase the solubility. From there, add more solute. Finally, cool to the original temperature.

Henry's Law

- The amount of dissolved gas in a liquid is proportional to its partial pressure above the liquid
- Why does a soda fizz when you open it?
    - Sodas contain dissolved carbon dioxide
    - Before opening, the gas above the drink in its container is almost pure carbon dioxide, at a pressure higher than atmospheric pressure
    - After the bottle is opened, this gas escapes, moving the partial pressure of carbon dioxide above the liquid to be much lower, resulting in degassing as the dissolved carbon dioxide comes out of the solution

Molarity and Molality

- Molarity (M): moles solute/L solution
    - Most commonly used
- Molality (m): moles solute/kg solvent
- % m/m: g solute per g solution ←

Dilution

stock ↓ diluted ↓

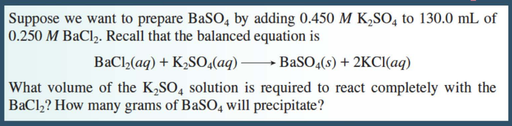
- M1V1=M2V2
- Example: How many mL of water is needed to dilute 85 mL of a stock NaOH (5 M) solution to 0.5 M?

$$(5M)(85mL)=(0.5M)(?\ mL)$$

↓

$$\frac{(5M)(85mL)}{(0.5M)} = \boxed{850\ mL}$$

Molar Concentration Example



Suppose we want to prepare $BaSO_4$ by adding 0.450 *M* K_2SO_4 to 130.0 mL of 0.250 *M* $BaCl_2$. Recall that the balanced equation is

$$BaCl_2(aq) + K_2SO_4(aq) \longrightarrow BaSO_4(s) + 2KCl(aq)$$

What volume of the K_2SO_4 solution is required to react completely with the $BaCl_2$? How many grams of $BaSO_4$ will precipitate?

$$\frac{0.250\text{ mol}}{\cancel{L}\ BaCl_2} \cdot \frac{1\cancel{L}}{1000\cancel{mL}} \cdot 130\cancel{mL} = 0.0325\text{ mol }BaCl_2$$

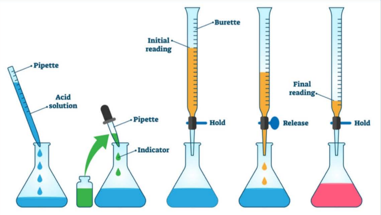
$$0.0325\text{ mol }BaCl_2 \cdot \frac{1\text{ mol }K_2SO_4}{1\text{ mol }BaCl_2} = 0.0325\text{ mol }K_2SO_4$$

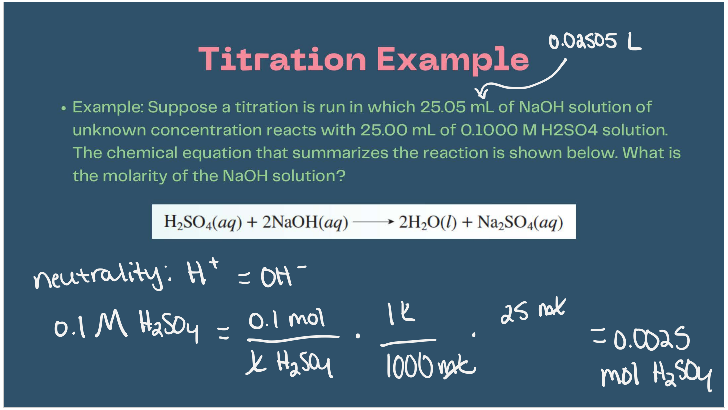
$$0.0325\ \cancel{\text{mol }K_2SO_4} \cdot \frac{1\ L}{0.450\ \cancel{\text{mol }K_2SO_4}} = \boxed{0.0722\ L\ K_2SO_4}$$

$$0.0325\text{ mol }BaCl_2 \cdot \frac{1\text{ mol }BaSO_4}{1\text{ mol }BaCl_2} = 0.0325\text{ mol }BaSO_4 \cdot \frac{233.4\ g}{1\text{ mol }BaSO_4} = \boxed{7.59\ g\ BaSO_4}$$

Titration

- The process of determining the concentration of one substance in solution
- by reacting it with a solution of another substance that has a known concentration





Titration Example

0.02505 L

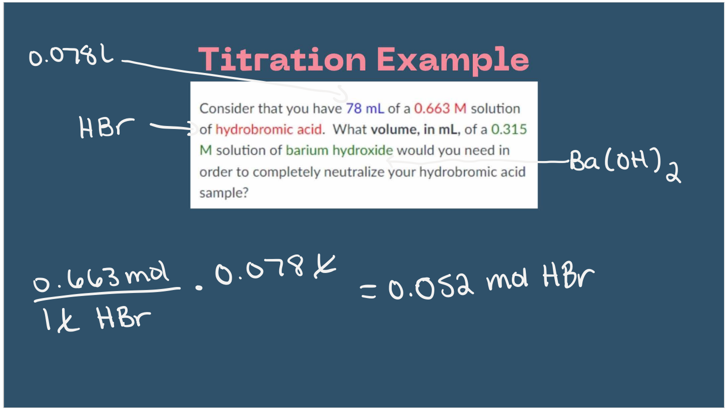
- Example: Suppose a titration is run in which 25.05 mL of NaOH solution of unknown concentration reacts with 25.00 mL of 0.1000 M H2SO4 solution. The chemical equation that summarizes the reaction is shown below. What is the molarity of the NaOH solution?

$$H_2SO_4(aq) + 2NaOH(aq) \longrightarrow 2H_2O(l) + Na_2SO_4(aq)$$

neutrality: $H^+ = OH^-$

$$0.1\ M\ H_2SO_4 = \frac{0.1\ \text{mol}}{\cancel{L}\ H_2SO_4} \cdot \frac{1\ \cancel{L}}{1000\ \cancel{mL}} \cdot 25\ \cancel{mL} = 0.0025\ \text{mol}\ H_2SO_4$$

$$0.0025\ \text{mol}\ H_2SO_4 \cdot \frac{2\ \text{mol}\ H^+}{\text{mol}\ H_2SO_4} \cdot \frac{1\ \text{mol}\ OH^-}{1\ \text{mol}\ H^+} = 0.005\ \text{mol}\ OH^- \cdot \frac{1\ \text{mol NaOH}}{1\ \text{mol}\ OH^-} \rightarrow \frac{0.005\ \text{mol NaOH}}{0.02505\ L} = \boxed{0.1996\ M}$$



Titration Example

0.078L

HBr

Ba(OH)$_2$

Consider that you have 78 mL of a 0.663 M solution of hydrobromic acid. What **volume, in mL**, of a 0.315 M solution of barium hydroxide would you need in order to completely neutralize your hydrobromic acid sample?

$$\frac{0.663\ \text{mol}}{1\ \cancel{L}\ HBr} \cdot 0.078\ \cancel{L} = 0.052\ \text{mol HBr}$$

$$0.052\ \text{mol HBr} \cdot \frac{1\ \text{mol}\ H^+}{1\ \text{mol HBr}} \cdot \frac{1\ \text{mol}\ OH^-}{1\ \text{mol}\ H^+} \cdot \frac{1\ \text{mol}\ Ba(OH)_2}{2\ \text{mol}\ OH^-} = 0.026\ \text{mol}\ Ba(OH)_2 \cdot \frac{1\ L\ \ Ba(OH)_2}{0.315\ \text{mol}\ Ba(OH)_2} \cdot \frac{1000\ mL}{1\ L} = \boxed{82.4\ mL}$$

$$2HBr + Ba(OH)_2 \rightarrow BaBr_2 + 2H_2O$$

$$0.052\ \text{mol HBr} \cdot \frac{1\ \text{mol}\ Ba(OH)_2}{2\ \text{mol HBr}} \cdot \frac{1\ L\ \ Ba(OH)_2}{0.315\ \text{mol}\ Ba(OH)_2} \cdot \frac{1000\ mL}{1\ L} = \boxed{82.5\ mL}$$