

# AP Chemistry Summary

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April 2021

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## 1 Notes

The list of things to memorize is not comprehensive; it includes/omits concepts based on what I knew before taking the class.

The PDF is not cut off. I just didn't have the time to finish the IMF section.

If you notice any typos or have any questions or suggestions, please email me at [michel.liao@systemgreen.org](mailto:michel.liao@systemgreen.org).

## 1.1 Things to Memorize

- Strong/weak acids/bases
- Phase changes
- The density of water is 1.0 g/mL
- Molecular Geometry
- Polyatomic ions
- Units ( $\text{nm} = 10^{-9}$ )
- Unit conversions
- Periodic table names
- Spectrophotometer
- Solubility rules
  - Compounds with an alkali metal cation ( $\text{Na}^+$ ,  $\text{Li}^+$ ,  $\text{K}^+$ , etc.) or an ammonium ( $\text{NH}_4^+$ ) are always soluble.
  - Compounds with a nitrate ( $\text{NO}_3^-$ ) anion are always soluble.
- Aluminum (Al), Zinc (Zn), and Silver (Ag) have +3, +2, +1 ion/oxidation charges/states, respectively (think of stair steps)
- Colors... solutions with permanganate are deep purple... when the manganese ion reduces to  $\text{Mn}^{2+}$ , it's colorless
- Microwave radiation is associated with changes in molecular rotational levels
- Infrared radiation is associated with changes in vibrational levels
- Ultraviolet/visible radiation is associated with transitions in electronic energy levels
- Intermediate: produced then consumed. Catalyst: consumed then produced
- When bonds are formed, energy is released. When bonds are broken, energy is absorbed.
- Enthalpy of formation is the change in energy that takes place when one mole of a compound is formed from its component pure elements

- Bond energy is the energy required to break a bond (sum of bonds broken minus bonds formed)
- Dissolving an ionic substance is done in three steps:
  1. Breaking the solute bonds, which has a positive  $\Delta H$ .
  2. Separating the solvent molecules (to make room for the ions), which has a positive  $\Delta H$ .
  3. Creating new attractions, which has a negative  $\Delta H$ .

Step 2 and 3 combined are called hydration energy, which is always negative since the 3rd step is always bigger than the 2nd step in magnitude.

- When liquid/solid phase are in equilibrium with gas phase, the pressure of the gas is equal to the vapor pressure
- If the vapor pressure is equal to atmospheric pressure, liquid boils.
- When a substance is changing phases, the temperature remains constant (you can't increase temp and phase change at the same time).
- When external pressure is increased, all the partial pressures will increase, so equilibrium will shift to the side with less gas molecules. If pressure is decrease, equilibrium will shift to the side with more gas molecules. (A change in volume is essentially a change in pressure.)
- When looking at how temperature affects equilibrium, add heat to the respective side (exo or endothermic), and treat it like a compound.
- When diluting with extra water, equilibrium shifts to the side with more aqueous species. When removing water, equilibrium shifts to the side with less aqueous species.
- Shifts caused by pressure/volume/concentration don't affect K because the products/reactants will establish the same ratio as earlier.
- Temperature changes **do** change K as well as shifting equilibrium.
- As acid molarity increases, percent ionization decreases (and vice versa).
- $E_{\text{cell}} = E_{\text{cathode}} - E_{\text{anode}}$

## 2 Random Notes

- A compound is diamagnetic if all its electrons are paired and paramagnetic if any of its electrons are unpaired. Paramagnetic compounds (and atoms) are attracted to magnetic fields while diamagnetic compounds (and atoms) are repelled from magnetic fields.

## 3 Atomic Structure and Properties

### 3.1 Moles & Gases, Percent Composition, and Ionization Energy

- At standard temperature and pressure ( $P = 1$  atm and  $T = 273$  K), one mole of gas occupies 22.4 liters.
- Percent composition is the percent by mass of each element that makes up a compound.
- Binding energy, or ionization energy, is the energy needed to remove an electron (it's always positive).
- When the valence shell is empty, the next electron must come from a shell that is much closer to the nucleus, making the ionization energy for that electron much larger than for the previous one.
- The second ionization energy is generally larger than the first.
  - After the first electron is removed, electron-electron repulsion decreases, and the remaining valence electrons get closer to the nucleus. This increases the attractive forces between the electrons and the nucleus, increasing ionization energy.

### 3.2 Empirical and Molecular Formulas

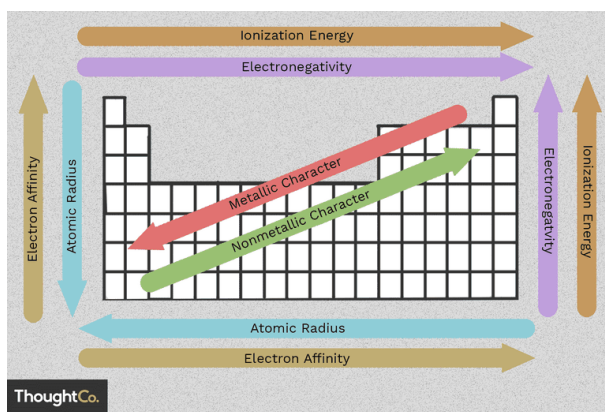
- The empirical formula is the simplest ratio of one element to another in a compound. Given percentages of a compound, assume we have 100g. of the compound. Then, find the moles of each compound and divide all of them by the smallest number of moles.
- The molecular formula is the actual formula for the substance. Picking up from the work above, we find the molar mass of the empirical formula. Then, we take the actual mass and divide it by the empirical formula's mass to get our scalar.

### 3.3 Configuration Rules

- The Aufbau Principle says that when building up the electron configuration of an atom, electrons are placed in orbitals, subshells, and shells in order of increasing energy.
- The Pauli Exclusion Principle says two electrons that share an orbital cannot have the same spin (one spins clockwise and the other counter-clockwise).
- Hund's Rule says that when an electron is added to a subshell, it will always occupy an empty orbital if one is available. Electrons will always occupy orbitals singly if possible.

### 3.4 Periodic Trends

1. Electrons are attracted to the protons in the nucleus.
    - The closer an electron is to the nucleus, the more attracted it is.
    - The more protons in a nucleus, the more strongly an electron is attracted.
  2. Electrons are repelled by other electrons. If electrons are between valence electrons and the nucleus, shielding happens.
  3. Completed shells (and to a lesser extent, subshells) are very stable. Atoms will add or subtract electrons to create complete shells if possible.
- Zinc always forms ions with a charge of +2 and silver always forms ions with a charge of +1 (think of Al, Zn, and Ag as steps of charges +3, +2, +1, respectively).



## 4 Molecular and Ionic Compound Structure and Properties

Bonding is the transfer or sharing of electrons. Atoms react in order to reach a more stable, lower-energy state. Atoms usually are at their most stable with eight electrons in their valence shells.

### 4.1 Ionic Bonds

- Ionic solids are held together by electrostatic attractions (giving electrons). They often occur between metals and nonmetals.
- The cation “gives” up some electrons to the anion, and the atoms are held together by the positive and negative charges on the ions.

- Ionic substances usually are solid at room temperature, and have high melting/boiling points. Their bonds are very strong.
- Two factors affect the melting/boiling points:
  1. A greater charge leads to greater bond energy (lattice energy), thus a higher melting/boiling point.
  2. Smaller ions have greater attraction, thus a higher melting/boiling point.
- Ionic solids don't conduct well because electrons are localized around an atom. Ionic liquids do conduct electricity because the *ions* are free to move (the electrons are still localized).

## 4.2 Metallic Bonds

- Use the sea of electron model when examining metals: The positively charged core (nucleus & core electrons) don't move, but the valence electrons are delocalized (free to move). This makes metals malleable and ductile.
- Alloys form from metals being bonded to one another. Typically, metals are melted, poured together, and cooled to create alloys.
  - An *interstitial alloy* forms when metal atoms with two vastly different radii combine (steel).
  - An *substitutional alloy* forms between atoms with similar radii (brass).

## 4.3 Covalent Bonds

- In covalent bonds, two atoms share electrons.

| <b>Bond type:</b> | <b>Single</b>          | <b>Double</b>                               | <b>Triple</b>                               |
|-------------------|------------------------|---|---|
| Bond designation: | One sigma ( $\sigma$ ) | One sigma ( $\sigma$ ) and one pi ( $\pi$ ) | One sigma ( $\sigma$ ) and two pi ( $\pi$ ) |
| Bond order:       | One                    | Two   | Three                                       |
| Bond length       | Longest                | Intermediate                                | Shortest                                    |
| Bond energy:      | Least                  | Intermediate                                | Greatest                                    |

- A covalent bond occurs when the potential energy of that bond is at the minimum possible level. If the atoms are too close, the nuclei repel each other. When they're too far apart, the nuclei cannot attract the electrons of the other atom.
- Network (Covalent) solids have very high melting and boiling points because the electrons are localized. Carbon (diamond/graphite) and silicon (quartz) are the most common examples of network solids.

#### 4.4 Lewis Dot Structures

- The least electronegative atom usually occupies the central position.
- It's okay if the central atom has more than 8 electrons, as long as it doesn't exceed 12.
- For resonance forms, you can find the bond order of a bond by adding up its bond order from all of the possible resonance forms and then dividing it by the number of resonance forms.
- Hydrogen is stable with 2 electrons in its outer shell (so is Helium, but Helium doesn't form bonds). Boron (B) is stable with 6 valence electrons.
- For molecules that have d subshells available, they can have more than 8 valence electrons ( $n = 3$  and above can have expanded octets).
  - This explains why noble gases can form bonds (extra electrons go into the d subshell).

#### 4.5 Formal Charge

- If there is more than one valid Lewis structure, we look at formal charges for a molecule.
  - Take the number of valence electrons for an atom and subtract the number of assigned electrons.
  - Bonds count as 1 and lone pairs count as 2 (each electron in a lone pair counts as 1, and  $1 + 1 = 2$ )
- The structure with fewer atoms with a formal charge (anything that isn't 0) is more likely. Formal charges should add up to 0 or whatever the overall charge is if the molecule is polyatomic ion.
  - If the formal charge of both structures is not 0, look for the one that is more consistent with electronegativities (e.g., F is more likely to have a formal charge of -1 than C).

#### 4.6 Molecular Geometry

- When we predict the geometries of molecules using the idea that molecules form shapes that keep their electron pairs far apart, we're using the valence shell electron-pair repulsion (VSEPR) model.
- We can find the hybridization of an atom using a point system: a lone pair, single bond, double bond, and triple bond all count as one point. If the total sums up to:  $2 = sp$ ,  $3 = sp^2$ ,  $4 = sp^3$ ,  $5 = sp^3d$ ,  $5 = sp^3d^2$ .
- Double bonds and triple bonds have slightly more repulsive strength, so they occupy slightly more space than a single bond
- Lone pairs have a little more repulsive strength than bonding pairs

## 5 Intermolecular Forces and Properties

### 5.1 Polarity

- A polar covalent bond is one where electrons are unequally shared (because of electronegativities).
- A dipole is an area of negative or positive charge.
- Generally, if there are lone pairs on the central atom, the molecule will be polar. The opposite is true, too.
  - The key exception is the square planar shape, which is frequently nonpolar.

### 5.2 Intermolecular Forces

- Intermolecular forces (IMFs) are the forces between molecules in a covalent substance.
- To change phases, a covalent substance's IMFs must be broken, but the bonds between the atoms stay in place. When ionic substances change phases, their bonds between ions actually break.
- Dipole-dipole forces occur when the positive end of one polar molecule is attracted to the negative end of another polar molecule.
- The larger the dipole moments, the higher melting/boiling points. Dipole-dipole attractions are weak and substances melt at low temperatures. Most dipole-dipole substances are gases or liquids at room temp.
- Hydrogen bonding is when a positively charged hydrogen is attached to a negatively charged end of another molecule containing N, O, or F (highly electronegative). These have high(er) melting/boiling points than substances only held together by other IMFs.
- London Dispersion Forces (LDFs) occur in all molecules when electrons temporarily gather on one side, forming a weak dipole. The more electrons, the greater the LDFs, and higher boiling/melting points. Only LDF substances are usually gas at room temp.

### 5.3 Bond Strength

- Ionic usually solid at room temp and melting requires breaking the bonds holding the lattice together.
- Covalent usually liquid at room temp and melting/boiling points of covalent usually lower than ionic.
- Metallic bonding tends to be very strong (esp. transition metals).



- Network covalent bonding is the strongest type of bonding.
- Vaporization is when a liquid molecule hits the surface with enough kinetic energy it transitions into the gas phase. The stronger the IMF of a molecule, the less likely molecules can escape the liquid, lower vapor pressure. Temperature and vapor pressure are directly proportional.

## 5.4 Solution Separation

- Like dissolves like (polar and ionic basically the same).
- The more ions in an ionic compound, the greater the conductivity when the ions are dissociated.
- Chromatography is the separation of a mixture by passing it in solution through a medium in which the components of the solution move at different rates.
- In paper chromatography, paper is used as the medium. The solute is put barely above the solvent line. Then, the paper touches the solvent and the solvent climbs up the paper. If the solvent is polar, then the most polar substance in the solution will travel the farthest. A major limitation is that if the solution had components that had no visible color, you can't see it on the filter paper.
- Column chromatography has a stationary substance in which a solution (analyte) is injected into. Another solution (eluent) is injected, and it travels down the column. The more alike the analyte is to the eluent, the farther it will travel.
- Distillation exploits the different boiling points of substances. If you heat the mixture up to the lowest boiling point, then one of the substances will boil but the others won't. The vapor will travel through a condenser that cools it back into liquid form. A major advantage is that the solutions don't need to be colored. A major disadvantage is when you don't know the boiling point of a substance.

## 5.5 Kinetic Molecular Theory

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$$KE = \frac{1}{2}mv^2$$

- Temperature and average kinetic energy are directly proportional. Average kinetic energy of a gas depends only on the absolute temperature, not the identity.
- The volume of an ideal gas particle is negligible compared to the volume the gas is in.

- There are no forces of attraction between gas molecules in an ideal gas.
- Gas molecules are in constant motion, colliding with one another and the container.

## 5.6 Maxwell-Boltzman Diagrams

- The higher the temperature, the larger the range of velocities.
- At the same temperature, the