

College Chemistry Second Semester Review Sheet

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Chapter 6: Thermochemistry

- I understand the law of conservation of energy and I can give examples of various energy types.
- I can distinguish between potential energy and kinetic energy and give examples for each.
- I can recognize and use the language of thermodynamics: system versus surroundings, open versus closed system, isolated system, endothermic versus exothermic processes.
- I can apply the first law of thermodynamics, $\Delta E = q + w$, and I know the conditions when q and w are positive or negative. See Table 1.

Table 1: Interpreting Signs for Heat and Work	
<i>Process</i>	<i>Sign</i>
<i>Heat released</i> by the system to the surroundings (exothermic process)	q is -
<i>Heat absorbed</i> by the system from the surroundings (endothermic process)	q is +
Work is done <i>by</i> the system <i>on</i> the surroundings	w is -
Work is done <i>on</i> the system <i>by</i> the surroundings	w is +

- I can explain the meaning of a state function and give examples of quantities that are state functions (ΔE and ΔH) as well as quantities that are not state functions (q and w).
- I can calculate work using $w = -P\Delta V$ and determine whether work is done *by* the system (gas expansion), *on* the system (gas compression), or whether *no work* is performed (piston doesn't move).
- I can compare and contrast constant-volume conditions ($\Delta E = q_v$) with constant-pressure conditions ($\Delta H = q_p$, $\Delta H = \Delta E + P\Delta V$).
- I understand what thermochemical equations are, and I can manipulate them using these guidelines:
 1. When writing thermochemical equations, we must specify the physical states of all reactants and products.
 2. If we multiply the coefficients of a thermochemical equation by a factor n , then the corresponding ΔH value must also change by the same factor.
 3. When a thermochemical equation is reversed so that the reactants become the products and vice versa, the magnitude of the corresponding ΔH value remains the same but the sign changes.
- I understand that 1 Food Calorie = 1kcal = 1000 calories and $R = 8.314 \text{ J/mole-K}$.
- I can compare and contrast the concepts of heat and temperature.
- I can use heat capacity, C , specific heat, s , latent heat of fusion, L_f , and latent heat of vaporization, L_v , to perform heat transfer calculations. Recall that the specific heat of liquid water is $4.184 \text{ J/g}^\circ\text{C}$.
- I can use $q = C\Delta T$, $q = ms\Delta T$, and $q = mL$ to calculate heat absorbed or heat released.
- I can compare and contrast constant-volume calorimeters (bomb calorimeters) with constant-pressure calorimeters (Styrofoam coffee cup calorimeters).
- I can correctly perform four types of calorimetry calculations:
 1. Combustion reaction done in a constant-volume "bomb" calorimeter.
 2. Hot metal dropped in water done in a constant-pressure Styrofoam cup calorimeter.
 3. Neutralization reaction done in a constant-pressure Styrofoam cup calorimeter.
 4. Phase change process (recall heating curve) done in a constant-pressure Styrofoam cup calorimeter.

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- I can write a balanced chemical equation that defines the standard molar enthalpy of formation, ΔH_f° , for a compound.

- I can directly calculate the enthalpy change for a reaction, ΔH_{rxn}° , using:

$$\Delta H_{rxn}^\circ = \sum n \Delta H_f^\circ(\text{products}) - \sum m \Delta H_f^\circ(\text{reactants})$$

- I can indirectly calculate the enthalpy change for a reaction, ΔH_{rxn}° , using Hess's Law.
- I can calculate the heat of solution from $\Delta H_{soln} = U + \Delta H_{hydr}$, in which U is the lattice energy (endothermic) and ΔH_{hydr} is the heat of hydration (exothermic).
- Due to the large, exothermic heat of dilution, I understand that concentrated acids are cautiously added to water rather than vice versa. This is a standard laboratory safety practice.

Chapter 7: Quantum Theory and the Electronic Structure of Atoms

- I understand electromagnetic spectrum concepts, wavelength (λ), frequency (ν), and amplitude.
- I can perform calculations using $c = \lambda\nu$ where $c = 2.998 \times 10^8$ m/s = speed of light.
- I can obtain the energy of a photon from either frequency or wavelength using $E = h\nu = \frac{hc}{\lambda}$. Recall that $h = 6.626 \times 10^{-34}$ Js = Planck's constant and $c = 2.998 \times 10^8$ m/s = speed of light.
- I can describe the photoelectric effect and how it supports energy being quantized at the atomic level.
- I can describe the Bohr model of the atom, its limitations, and how the model accounts for the emission line spectra of excited atoms.
- I understand that in the Bohr model of the hydrogen atom, electrons can occupy only certain energy levels, each with an energy given by $E_n = -R_H \left(\frac{1}{n^2} \right)$, where $R_H = 2.18 \times 10^{-18}$ J = Rydberg constant and n is the principal quantum number (an integer ≥ 1).
- I understand that if an electron moves from an initial state, n_i , to a final state, n_f , the energy absorbed or emitted is equal to the difference in energy between the two states. I can calculate this energy using $\Delta E = h\nu = R_H \left(\frac{1}{n_i^2} - \frac{1}{n_f^2} \right)$.
- For the modern view of the atom, I can calculate the wavelength of an electron or any subatomic particle using de Broglie's equation $\lambda = \frac{h}{mu}$ where m = mass and u = velocity for the particle.
- I can describe the allowed energy states of the electron in an atom using the quantum numbers in Table 2. I can give the name, allowed values, and orbital property for each quantum number.

Table 2: The Quantum Numbers			
<i>Quantum Number</i>	<i>Symbol</i>	<i>Allowed Values</i>	<i>Orbital Property</i>
Principal quantum #:	n	1, 2, 3, ...	Orbital size
Angular momentum quantum #:	ℓ	0, 1, 2, ... $n-1$	Orbital shape
Magnetic quantum #:	m_ℓ	0, ± 1 , ± 2 , ... $\pm \ell$	Orbital orientation
Spin quantum #:	m_s	$+\frac{1}{2}$, $-\frac{1}{2}$	Spin pairing of electrons in orbitals

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- I can describe subshell labels and corresponding orbital shapes for ℓ values 0 to 3. See Table 3.

ℓ -Value	Subshell	Orbital Shape	# of Orbitals	# of Electrons
0	s	Sphere	1	2
1	p	Dumbbell	3	6
2	d	Four-leaf clover	5	10
3	f	Two four-leaf clovers	7	14

- I can give the number of electrons in a particular shell ($2n^2$), subshell (see table above), and orbital (2 electrons).
- I understand the Heisenberg uncertainty principle: It is impossible to know simultaneously both the momentum and the position of a particle with certainty.
- I understand that the position of an electron is not known with certainty; only the *probability* of the electron being within a given region of space (an “orbital”) can be calculated.
- I can apply the Pauli exclusion principle: No two electrons in an atom can have the same set of four quantum numbers.
- I understand that electrons are assigned to the subshells of an atom in order of increasing subshell energy. In the H atom, the subshell energies increase with increasing n , but in a many-electron atom, the energies depend on both n and ℓ .
- I can use the Aufbau principle to help write electron configurations for atoms of particular elements.
- When writing orbital box diagrams, I can apply Hund’s rule: The most stable arrangement of electrons in subshells is the one with the greatest number of parallel spins (in other words, electrons occupy all the orbitals of a given subshell singly before electron pairing begins).
- I can write orbital box diagrams and classify substances as paramagnetic (has unpaired electrons) or diamagnetic (does not have unpaired electrons).

Chapter 8: Periodic Relationships Among the Elements

- I can explain Mendeleev’s and Moseley’s contributions to the shape of the modern periodic table.
- I can identify the alkali metals, alkaline earth metals, transition metals, halogens, noble gases, lanthanides, and actinides on the periodic table. I can also distinguish between the metals, nonmetals, and metalloids (or semimetals).
- I can identify the s-block, p-block, d-block, and f-block elements on the periodic table.
- I can use the periodic table as a guide to determine electron configurations for atoms and ions of particular elements. Be able to express the electron configurations using an orbital box notation or a spectroscopic notation.
- I can express electron configurations using noble gas notation.
- I can predict how properties of atoms like atomic radius, ionization energy (IE), and electron affinity (EA) change within a group or across a period of the periodic table.
- I can explain the role that atom and ion size, ionization energy, and electron affinity play in the chemistry of the elements.
- I can predict how atomic radius changes when atoms form ions. In general, when neutral atoms form cations, they decrease in size; when neutral atoms form anions, they increase in size.
- I can compare ionic radii by classifying ions into three useful categories as shown in Table 4.

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Table 4: Comparing Ionic Radii for Different Types of Ions	
<i>Category</i>	<i>Ionic Radii Comparison</i>
1. Isoelectronic ions:	Ions carrying a greater negative charge are always larger.
2. Ions that carry the same charges and are generated from atoms of the same group in the periodic table:	Ions from atoms having a greater atomic number are always larger.
3. Ions that carry different charges but are generated from the same atom:	Ions having a smaller positive charge are always larger.

- I can calculate ionization energies for a one-electron, hydrogenlike ion using

$$E_n = -(2.18 \times 10^{-18} \text{ J})Z^2 \left(\frac{1}{n^2} \right)$$

where n is the principal quantum number and Z is the atomic number of the element.

- I understand that metal oxides tend to be basic and nonmetal oxides tend to be acidic.
- I can write balanced chemical equations for the formation of simple bases and acids when metal oxides and nonmetal oxides, respectively, react with water.

Chapter 9: Chemical Bonding I: Basic Concepts

- I can explain the difference between core electrons and valence electrons.
- I can write Lewis dot symbols for the elements and show the gain or loss of electrons to form ionic compounds.
- I understand the concept of lattice energy and how it is a quantitative measure of stability for any ionic solid. As lattice energy increases, the stability of the ionic solid increases.
- I can estimate changes in lattice energy from the potential energy form of Coulomb's Law,

$$E = k \frac{Q_{\text{Cation}} Q_{\text{Anion}}}{r}$$

in which E is lattice energy, k is a proportionality constant, Q_{Cation} is the

charge on the cation, Q_{Anion} is the charge on the anion, and r is the distance between the centers of the ions.

- I can determine lattice energies using a Born-Haber cycle, which is a Hess's Law process using ionization energies, electron affinities, and enthalpies for other atomic and molecular properties.
- I can compare and contrast ionic and molecular compounds. See Table 5.
- I can describe the basic forms of chemical bonding—ionic and covalent—and the differences between them.
- I can compare and contrast the properties of ionic and covalent (molecular) compounds.
- I can predict trends in bond length when comparing carbon-carbon and carbon-oxygen single, double, and triple bonds.

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Table 5: Comparing Ionic and Molecular Compounds		
	<i>Ionic Compounds</i>	<i>Molecular Compounds</i>
Bonding Type:	Ionic Bonding	Covalent Bonding
In this type of bonding, electrons are _____:	Transferred	Shared
Type(s) of Elements Involved:	Metal + Nonmetal Elements	Nonmetal Elements
Comparison of electronegativity differences:	Larger	Smaller
Comparison of Properties:		
a. Melting and boiling points:	a. Higher	a. Lower
b. Hardness:	b. Harder	b. Softer
c. Conduction of electricity:	c. When molten or dissolved in water, ionic compounds tend to conduct electricity.	c. Molecular compounds do not conduct electricity.

- I can use electronegativity trends in the periodic table to compare atoms of various elements.
- I can use electronegativity differences to classify bonds as nonpolar covalent, polar covalent, and ionic. See Table 6. I also know the limitations for the given table.

Table 6: Classifying Bonds Using Electronegativity Differences	
<i>Electronegativity Difference</i>	<i>Bond Type</i>
0 - 0.2	Nonpolar covalent bond
0.3 - 1.9	Polar covalent bond
≥ 2.0	Ionic bond

- I can apply the octet rule to help write Lewis structures for covalent compounds and polyatomic ions.
- I remember that hydrogen violates the octet rule and can never have more than two electrons around it in a Lewis structure.
- I can count the number of bonding and nonbonding electrons around any atom in a Lewis structure, and I recognize that nonbonding pairs are sometimes called “lone pairs” of electrons.
- I can calculate formal charges for atoms in any Lewis structure and use them to determine how reasonable the structure is. See Table 7.

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Table 7: Formal Charge Guidelines for Drawing Lewis Structures

- (1) For molecules, a Lewis structure in which there are no formal charges is preferable to one in which formal charges are present.
- (2) Lewis structures with large formal charges (+2, +3, and/or -2, -3, and so on) are less plausible than those with small formal charges.
- (3) Among Lewis structures having similar distributions of formal charges, the most plausible structure is the one in which negative formal charges are placed on the more electronegative atoms.

- I can write Lewis structures for molecules that do not obey the octet rule (molecules having incomplete octets, expanded octets, or an odd number of electrons).
- I can explain the difference between a normal covalent bond and a coordinate covalent bond.
- I understand what resonance means and I can write resonance structures for common polyatomic ions.
- I can calculate bond order using $Bond\ Order = \left(\frac{Number\ of\ shared\ pairs\ of\ electrons}{Number\ of\ regions\ of\ electrons} \right)$ for a given molecule.
- I can use bond enthalpies to distinguish between the relative strengths of various covalent bonds.
- I can use bond enthalpies to calculate heats of reaction, ΔH_{rxn}° using

$$\Delta H_{rxn}^{\circ} = \Sigma BE(\text{reactants}) - \Sigma BE(\text{products}) = \text{total energy input} - \text{total energy released}$$
 where BE is the various bond enthalpies for the reactants and products.

Chapter 10: Chemical Bonding II: Molecular Geometry and Hybridization of Atomic Orbitals

- I can predict the shape of covalent molecules and polyatomic ions using Valence Shell Electron Pair Repulsion (VSEPR) Theory. I can also name the electron-pair geometry and the molecular geometry.

1. “Electron-pair geometry” refers to the structural arrangement of the **electron pairs**:

<u>Number of Regions of Electron Pairs</u>	<u>Name of Electron-pair Geometry</u>	<u>Bond Angle(s)</u>	<u>Hybridization</u>
2	linear	180°	sp
3	trigonal planar	120°	sp ²
4	tetrahedral	109.5°	sp ³
5	trigonal bipyramidal	90°, 120°	sp ³ d
6	octahedral	90°	sp ³ d ²

2. “Molecular Geometry” refers to the structural arrangement of the **atoms**:

<u>Structural Type</u>	<u>Molecular Geometry</u>	<u>Structural Type</u>	<u>Molecular Geometry</u>
AB ₂	linear	AB ₅	trigonal bipyramidal
AB ₃	trigonal planar	AB ₄ E	see-saw
AB ₂ E	bent	AB ₃ E ₂	T-shaped
AB ₄	tetrahedral	AB ₂ E ₃	linear
AB ₃ E	trigonal pyramidal	AB ₆	octahedral
AB ₂ E ₂	bent	AB ₅ E	square pyramidal
		AB ₄ E ₂	square planar

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- Knowing the electron-pair geometry, I can determine the corresponding orbital hybridization and bond angle(s) present.
- I can use electronegativity values to determine bond polarity.
- I can combine knowledge of bond polarity and molecular geometry to predict molecular polarity.
- With regard to covalent bonding, I can describe the main features of valence bond theory and molecular orbital theory.
- I recognize that the premise for valence bond theory is that bonding results from the overlap of atomic orbitals. This orbital overlap results in electrons being concentrated (or localized) between two atoms.
- I can use valence bond theory to describe orbital hybridization, sigma (σ) and pi (π) bonding, and potential energy curves for covalent bond formation.
- I can predict trends in bond length, bond order, and bond enthalpy for carbon-carbon single, double, and triple bonds. I can also describe the sigma (σ) and pi (π) bonding present for each.
- I can distinguish how sigma (σ) and pi (π) bonds arise. For σ bonding, orbitals overlap in a head-to-head fashion, concentrating electrons along the bond axis. Sideways overlap of p atomic orbitals results in π bond formation, with electrons above and below the molecular plane.
- I can understand molecular orbital (MO) theory. In MO theory, atomic orbitals combine to form bonding and antibonding molecular orbitals that can be delocalized over several atoms.
- With regard to MO theory, I can answer questions about electron configurations, magnetism, bond order, bond length, and bond stability.
- Using MO theory, I can calculate bond order from

$$\text{Bond Order} = \left(\frac{1}{2} \right) (\text{Number of Bonding Electrons} - \text{Number of Antibonding Electrons}).$$

- I can describe electron delocalization in benzene and in common polyatomic ions like CO_3^{2-} .

Chapter 11: Intermolecular Forces, Liquids, and Solids

- I can use the Kinetic Molecular Theory of Matter to compare and contrast the physical properties of gases, liquids, and solids (see Table 11.1 in text). Focus on volume, shape, compressibility, relative density, and molecular motion.
- I can compare and contrast the intermolecular forces of attraction (IFA or IMF; sometimes called “van der Waals forces”) for molecular compounds. See Table 8.

Table 8: Intermolecular Forces of Attraction (“van der Waals forces”)			
<i>Type</i>	<i>Strength</i>	<i>Attraction</i>	<i>Examples</i>
Hydrogen bonding	Strongest	“H” on one molecule is attracted to “F,” “O,” or “N” on an adjacent molecule.	H ₂ O, NH ₃ , HF, base pairs in DNA
Dipole-dipole forces	Intermediate	Polar molecules attract each other like tiny magnets. (Larger dipole moments result in greater interactions.)	HCl, H ₂ S
London dispersion forces	Weakest	Induced dipoles attract each other for a fraction of a second; works best with atoms and molecules having lots of electrons, which makes them more polarizable. All covalent compounds have London dispersion forces.	CH ₄ , CCl ₄

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- I understand that as molar mass increases, melting points of similar nonpolar compounds increase due to a greater number of electrons and the resulting ease in polarizability (see Table 11.2 in text).
- I can use intermolecular forces of attraction to explain relative melting points, boiling points, surface tensions, viscosities, specific heats, and vapor pressures for a given series of molecular compounds.
- I can compare and contrast intramolecular forces of attraction with intermolecular forces of attraction for ionic, covalent, molecular, and metallic materials. See Table 9.

Table 9: Comparing Intramolecular and Intermolecular Forces of Attraction		
<i>Intramolecular Force</i>	<i>Intermolecular Force(s)</i>	<i>Examples</i>
Ionic Bond	Electrostatic force	NaCl, BaF ₂
Covalent Bonds forming a Network Solid	Covalent Bond	C(diamond), SiO ₂
Covalent Bond	Hydrogen bonding Dipole-dipole forces London dispersion forces	See examples in previous table.
Metallic Bond	(“sea of electrons”)	Ag, Na

- I can distinguish between surface tension, capillary action, cohesion, and adhesion.
- I can explain why water has “surprising” physical properties
 - Water is polar because it has polar bonds and bent geometry. It is also capable of hydrogen bonding, which resulting in several unusual properties.
 - Water has a higher melting point and boiling point than expected.
 - Water has considerable surface tension.
 - Water has a high capacity to absorb heat.
 - Ice is less dense than liquid water and will therefore float.
 - In addition to these properties, polar and ionic substances will dissolve in water.
- I can compare and contrast the general properties and the forces holding units together for ionic, covalent, molecular, and metallic crystals. See Table 10.
- I recognize that for ionic compounds like NaCl, the crystalline solid is generally a poor conductor of electricity, but its melt and an aqueous solution containing the substance both conduct electricity.
- I can draw and explain phase diagrams. I can explain the significance of the triple point and the critical point.
- I can draw and explain heating curves. I can perform related calculations starting at any point on the curve and ending at any other point on the curve.
- I understand how molar heats of fusion, vaporization, and sublimation are related to each other: $\Delta H_{fus} + \Delta H_{vap} = \Delta H_{sub}$
- I can perform calculations with the Clausius-Clapeyron equation:

$$\ln\left(\frac{P_1}{P_2}\right) = \frac{\Delta H_{vap}}{R} \left(\frac{1}{T_2} - \frac{1}{T_1}\right)$$

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Table 10: Crystal Types and General Properties			
<i>Type of Crystal</i>	<i>Force(s) Holding the Units Together</i>	<i>General Properties</i>	<i>Examples</i>
Ionic	Electrostatic attraction	Hard, brittle, high melting point, poor conductor of heat and electricity	NaCl, LiF, MgO, CaCO ₃
Covalent	Covalent bond	Hard, high melting point, poor conductor of heat and electricity	C (diamond), SiO ₂ (quartz)
Molecular	Hydrogen bonds, Dipole-dipole forces, London dispersion forces	Soft, low melting point, poor conductor of heat and electricity	Ar, CO ₂ , I ₂ , H ₂ O, C ₁₂ H ₂₂ O ₁₁ (sucrose)
Metallic	Metallic bond	Soft to hard, low to high melting point, good conductor of heat and electricity	All metallic elements; for example, Na, Mg, Fe, Cu

Chapter 23: Nuclear Chemistry

- I can explain the relationship between rest energy and mass, $E_o = mc^2$.
- I can explain how the nuclear force (strong force) holds the nucleus together despite it being composed of protons and neutrons, all of which have a positive charge or no charge, respectively.
- I can distinguish between alpha, beta, and gamma radiation, and I can identify the associated particles.
- I can write and balance equations for nuclear reactions.
- I can explain what happens in a uranium-238 decay series, how it produces radon gas, and how radon gas decomposes in a person's lungs to help make it the second leading cause of lung cancer behind smoking.
- I can determine half-life from $T_{1/2} = \frac{0.693}{\lambda}$, and I can explain how half-life helps to measure nuclear decay for an isotope of a particular element.
- I can distinguish between nuclear fission and nuclear fusion, and I can give examples for each.
- I can intelligently discuss nuclear energy concerns in everyday life:
 - What health hazards are associated with radioactivity?
 - Can a nuclear power plant undergo a nuclear explosion?
 - What are the safeguards against "meltdown"?
 - How will society dispose of waste from nuclear power plants?
 - Can nuclear fuel and nuclear waste be diverted to make weapons of mass destruction that terrorists may use?