

Things to Know – Chem 1A

Chapter 4

1. Definitions of oxidation and reduction.
2. Be able to find the oxidation number of any atom (in an element or compound)
3. Given a reaction, determine what is oxidized and what is reduced (first find oxidation numbers).
4. If an element appears in a reaction, the reaction must be a redox reaction.
5. Understanding the activity series and how to determine relative activities:
 $A + BC \rightarrow B + AC$ If this happens, A is more active than B. Given the activity series, be able to predict whether a particular single-replacement reaction will occur.
6. Molarity = # moles solute / L solution. Be able to calculate molarity or use it as a conversion factor.
7. Moles of solute = $M \times V$ of a solution. V must be in liters here.
8. Concentration of ions: 0.1 M Na_2SO_4 (aq) is 0.2 M Na^+ and 0.1 M SO_4^{2-} .
9. Describe how to make a solution of a given molarity. (Use a volumetric flask. Describe the process in words.)
10. Dilutions – you can use $M_1V_1 = M_2V_2$, where V_2 is the total final volume of solution ($V_2 = V_1 + \text{water added}$). Why? Because # moles of solute doesn't change.
11. Stoichiometry problems involving solutions
12. Titrations – no limiting reactant. Need a balanced equation. Start with the substance you can find moles of. Convert to moles of the other substance. Finish the problem: could find molarity, volume, mass %, molar mass, etc. **DO NOT** use $M_1V_1 = M_2V_2$. Why not?
13. Limiting reactant problem involving solutions: If you mix two solutions and a precipitate forms, determine all ion concentrations after the reaction. Approach: write the net ionic equation for the reaction. Find the initial number of moles of each ion. Understand that the spectator ions don't react. Of the ions that react, find which is the LR. Determine how much of the excess ion remains after the reaction. Then find concentrations of all remaining ions by dividing by the total volume of the mixed solutions.

Chapter 10

1. Properties of gases
2. Pressure – what is it? How is it measured? How does atmospheric pressure vary? What are gas molecules doing to exert pressure?
3. How does a barometer work? What does it measure? Why use Hg as the liquid?

4. How does an open-end manometer work? What about a closed-end manometer? What does each one measure?
5. Boyle's Law: $P_1V_1 = P_2V_2$ As pressure increases, volume decreases. (Used if n and T are constant.) **Why?** Explain at the molecular level.
6. Charles' Law: as T increases, volume also increases. V is proportional to temp in Kelvin. (Used if n and P are constant.) **Why?** Explain at the molecular level.

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

7. Gay-Lussac's Law: as T increases, P increases (assuming n and V are constant). **Why?** Explain at the molecular level.

$$\frac{P_1}{T_1} = \frac{P_2}{T_2}$$

8. Combined gas law (used when n is constant but P, V and T are changing)

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

9. Avogadro's law: volume is proportional to moles of gas (assume P and T are constant). **Why?** Explain at the molecular level.

$$\frac{V_1}{n_1} = \frac{V_2}{n_2}$$

10. What is STP? What is the molar volume of an ideal gas at STP? (22.4 L/mole : can use it as a conversion factor.)
11. Reacting gases: if measured at the same T and P, volume is proportional to moles. (Can use volume ratios instead of mole ratios – this is an important shortcut!)
12. Ideal gas law $PV = nRT$, $R = 0.08206 \text{ L}\cdot\text{atm}/\text{K}\cdot\text{mol}$ Units used must match the units of R.
13. Be able to derive any of the other gas laws from $PV = nRT$ (put constants on one side, variables on the other)
14. Given the density of a gas at a specific temperature and pressure, calculate MM of a gas (# g/# mol) Approach: assume 1 L. Find the number of moles of gas in 1 L, then divide # g/L by #mol/L.
15. Given the molar mass of a gas, P, and T, calculate density. Approach: find the number of moles in 1 L, then convert to g.
16. Stoichiometry problems involving gases, including limiting reactant problems
17. Gas mixtures - partial pressures $P_{\text{total}} = P_A + P_B + P_C + \dots$,
18. $P_{\text{total}} = n_{\text{total}}RT/V$ $P_A = n_ART/V$ $P_A = X_AP_{\text{total}}$
19. Collect a gas over water – subtract vapor pressure of water from total pressure – why?
20. In Lab 13 calculation, why did you have to subtract an extra correction factor from P_{total} to get P_{H_2} ?
21. Kinetic-Molecular theory – **how it explains behavior of gases**

22. Be able to explain any of the gas laws at the molecular level (explain what the molecules are doing and how that gives rise to the observable properties). Be able to explain these relationships: P and V, V and T, P and T, n and V, n and P.
23. Velocity distribution of gas molecules (Boltzmann distribution) – low vs. high temp. (What does the graph look like? Are all the molecules in a sample traveling the same speed?) Be able to sketch this graph, label the sketch, and explain what it means. The x axis could be either velocities or kinetic energies of gas molecules.
24. Using the equation below: what units are used? What is u_{rms} ?

$$u_{rms} = \sqrt{\frac{3RT}{M}}$$
25. Diffusion or effusion problems – comparing rates, times, # moles, etc. – reason it out! Make sure your answer makes sense. Molar mass ratio goes under the square root symbol. Practice these types of problems.
26. Nonideal gases – under what conditions do gases exhibit nonideal behavior? Explain the underlying reason for each.
27. Van der Waals equation – what does it take into account?

Chapter 5

1. Energy, work, conservation of energy. Kinetic vs. potential energy. Types of potential energy
2. Units of energy
3. system, surroundings, closed system
4. heat, thermal energy (molecular level explanation for things next to each other ending up at the same temp.)
5. First law of Thermo: $\Delta E = q + w$. What is included in E?
6. Sign conventions for q and w
7. Endothermic vs. exothermic
8. State function vs. path function. What is the significance of state functions?
9. Know how to calculate w for a given reaction. $w = -P\Delta V = -\Delta n_{gas}RT$
10. Enthalpy (ΔH) = heat transferred at constant pressure. Heat transferred at constant volume = ΔE . $\Delta E = \Delta H - P\Delta V$ or $\Delta E = \Delta H + w$
11. Thermochemical equations – understand the implied amounts (kJ per x moles)
12. Stoichiometry using ΔH – how much energy is given off per x g of reactant? Example: $2 \text{ H}_2 + \text{O}_2 \rightarrow 2 \text{ H}_2\text{O}$ $\Delta H = -572 \text{ kJ}$ means:

$\frac{-572 \text{ kJ}}{2 \text{ mol H}_2}$	$\frac{-572 \text{ kJ}}{1 \text{ mol O}_2}$	$\frac{-572 \text{ kJ}}{2 \text{ mol H}_2\text{O}}$
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13. Specific heat (C) vs. heat capacity (also C) (units?) They are different!

14. What does it mean if a substance has a high specific heat? What about a low specific heat?
15. Use $q = Cm\Delta T$ or $q = C\Delta T$ for temperature changes.
16. Coffee-cup calorimetry – constant P (open container) $q_P = \Delta H$
 $q_{\text{calorimeter}} = q_{\text{solution}}$ (Usually the heat capacity of the styrofoam cup itself is ignored.) We often have to assume that the solution in the calorimeter has a density and heat capacity close to those of water, since the solution consists of mostly water.
17. Bomb calorimetry – rigid container, so $\Delta V = 0$, so $q_V = \Delta E$.
 $q_{\text{calorimeter}} = (q_{\text{water}} + q_{\text{bomb}})$ Sometimes, the heat capacity of the entire calorimeter is given with the water already included. In that case,
 $q_{\text{calorimeter}} = C_{\text{calorimeter}}\Delta T$.
18. For both of the above types of calorimetry, $q_{\text{calorimeter}} = -q_{\text{reaction}}$! In order to get ΔH or ΔE , divide by the number of moles of the limiting reactant. Then check the stoichiometry of the reaction – adjust for coefficients.
19. Manipulating thermochemical equations (reverse a rxn, multiply a rxn by N, add reactions) – what happens to ΔH in each case?
20. Hess's Law: given several reactions and their ΔH values, determine the ΔH value for a new reaction by rearranging and adding up the given reactions.
21. Heat of formation – know how it's defined, be able to write the corresponding reaction.
22. Given ΔH°_f values, calculate ΔH_{rxn} (products – reactants)
23. OR Given ΔH°_f values and ΔH_{rxn} , calculate a missing ΔH°_f value. (Don't forget: for elements, the $\Delta H^\circ_f = 0$. Why?)
24. Phase changes - constant temp (Why? Be able to explain at the molecular level). Be able to draw a heating curve.
25. Calculating total energy involved in a process that has phase changes and temperature changes (calculate them separately, then add)
26. ΔH_{vap} , ΔH_{fus} – What are they? Which one is higher? Why?
27. Which phase changes have a positive q ? Which have a negative q ?
28. Lab 15 calculation!
29. Heat involved in phase changes practice worksheet!
30. Heat is also involved in making a solution and diluting. (given calorimetry data, calculate ΔH of solution.) ex: $\text{NaCl}_{(s)} \rightarrow \text{Na}^+_{(aq)} + \text{Cl}^-_{(aq)}$ Be able to calculate ΔH for these processes too.

Chapter 6

1. Wave properties: wavelength, amplitude, velocity, frequency
2. Types of electromagnetic radiation – which ones are high energy? Which are low energy? Which color of visible light has the highest frequency and which color has the lowest frequency? What is the order of energy of the different types of electromagnetic radiation?
3. $c = \nu\lambda$
4. What is a node?

5. What is a standing wave? What is the significance of a standing wave?
6. What is quantization?
7. Failures of classical physics: blackbody radiation, photoelectric effect. What new assumption was used to resolve these issues?
8. General relationship between E , ν , and λ .
9. Calculate E per photon or E per mole of photons, calculate wavelength from E , etc.
10. Bohr model for H: allowed energy levels for the electron are:

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

$$\Delta E_{\text{e-}} = -2.18 \times 10^{-18} \text{J} \left(\frac{1}{n_f^2} - \frac{1}{n_i^2} \right)$$

11. What are possible values of n in the above equation?
12. $\Delta E_{\text{electron}} = E_{\text{photon}}$
13. Ground state vs. excited state
14. Energy levels get closer together as you go up.
15. emission vs. absorption (sign of ΔE)
16. Calculate E or λ of a certain transition (for H) – use above equation. Which transitions occur in the visible region?
17. Why are only certain wavelengths of light emitted/absorbed?
18. Understand the connection between what the electron is doing and the light wavelength involved.
19. Spectrum of H (for other atoms, the spectrum is more complicated.)
20. Why do different atoms emit different colors of light?
21. De Broglie equation: used for matter, not light

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

22. Uncertainty principle: we can't know where the electron is, but we can get info on the probability of finding it in a certain region.
23. Ideas behind quantum mechanics
24. What is a wavefunction? What is the physical significance of ψ^2 ?
25. Quantum numbers: n , l , m_l , m_s : possible values, physical meaning for each
26. Know the general shape of s, p, and d orbitals
27. Be able to sketch different types of orbitals – include nodes. $n-l-1$ radial/spherical nodes, l nodal planes/angular nodes
28. For the H atom, the energy of the orbitals depends only on n .
29. For “many-electron” atoms, the energy depends on n and l .
30. Electron spin is quantized: $m_s = +1/2$ or $-1/2$.
31. Pauli exclusion principle, Aufbau principle, Hund's rule (what are they?)
31. Ground state vs. excited state vs. impossible configurations
32. Order of energy: 1s, 2s, 2p, 3s, 3p, 4s, 3d, 4p, 5s, 4d, 5p, etc. Know this order of energy (use the periodic table to remember)
33. Be able to write the electron configuration for any atom.
34. Arrow/orbital diagrams, spectroscopic notation, noble gas notation

35. Paramagnetic, diamagnetic, ferromagnetic – what is the difference? How can you tell the difference experimentally?
36. Exceptions to filling order: Cr, Cu, others.
37. Elements in the same group have similar chemical properties. Why?