Review Guide: Exam 3, Chemistry 115

**Chapter 8 – Chemical Equations**

Be able to predict chemical reactions

Use Activity Series and Solubility Rules.

**CH 9: Chemical Quantities in Reactions**

**Stoichiometry:** Use mole-to-mole ratios to relate and calculate amounts of reactants and/or products in a chemical reaction

**Limiting Reagent Problems**

• Calculate the mass or volume of product that can be made using the given amounts of each reactant and the balanced chemical equation.

* Solve using the comparison-of-moles or comparison-of-mass method (i.e., determine if there is enough of one reactant to react completely with the other)
  + Indicate the limiting reactant (or reagent) and the reactant(s) in excess.
  + Calculate the amount of reactant in excess that remains after the reaction.
  + Use conservation of mass to calculate the mass of reactant in excess that remains after the reaction or the mass of product(s) that form(s).
  + Use stoichiometry to calculate the energy absorbed or released by a reaction for given amount of reactant or product.

**Yields of Reactions**

* + theoretical yield: amount of product predicted using the balanced equation when limiting reagent is used up (can be calculated)
  + actual yield: amount of product one actually obtains (generally given in the problem)

**Chapter 10 – Modern Atomic Theory**

**wavelength** (λ): distance between peaks

**frequency** (ν): number of crests passing by a

given point in 1 s; given in 1s =hertz (Hz)

Wavelength (λ) is inversely related to

frequency (ν) and energy (E):

– As λ↑ → ν↓, Ε↓ or As λ↓ → ν↑,Ε↑

**Electromagnetic Spectrum:**

* Continuum of radiant energy
* Gamma (γ) rays to radio waves
* The visible spectrum makes up a small portion
* Red light at 700 nm is lower in energy than blue light at 400 nm

**Bohr Theory of the Atom**

* Electrons move in quantized orbits called “energy levels” around nucleus
* **ground state:** electron(s) fill the lowest energy level(s) before any in higher level(s)
* **excited state:** electron(s) in higher energy level(s) before lower levels are full
* When an atom absorbs energy, e− jumps from lower energy to higher energy level.
* When an e− drops from a higher energy to lower energy level, it releases energy, sometimes as light → atomic emission spectra

**Schrödinger’s** equation → probability of finding the electron in a given region in space

→ probability density = electron cloud

→ “shape” of atomic orbitals

**Atomic Orbital Shapes**

* reflect the “probability density” for an electron in a given orbital
* As **principal energy level** (**n**=1, 2, 3,…) increases, the orbital size increases.
* Energy levels divided into sublevels (s, p, d, f)
* Know shapes and number of s, p, and d orbitals (e.g. one s, three p, five d orbitals).
* Be able to write ground state **electron**
* **configurations for any neutral atom or ion**
  + Write using full notation and core notation
* (Noble Gas abbreviation)
* Recognize extra stability gained with filled and half-filled d orbitals (Cr, Mo, Cu, Ag)
* Account for electrons gained or lost for ions
* Know Representative Elements usually form ions that are **isoelectronic** with a Noble Gas Know definitions for atomic radius, metallic character, and ionization energy.

**Chapter 11 – Chemical Bonds**

Know **Periodic Trends for**

• **Atomic radius:**

* Increases down a group
* Decreases left to right across a period because effective nuclear charge increases
* **Metallic Character:** same trends as atomic radius
* **Ionization Energy (IE):**
* Decreases down a group
* Increase left to right across a period
* Opposite trend as atomic radius since
* IE ↓ as atomic radius ↑

**core electrons:** innermost electrons in filled

shells

**valence electrons:** outermost electrons

* The group number for each element is equal to its number of valence electrons
* Recognize that only valence electrons are gained, lost, or shared during chemical reactions.
* Draw **Electron Dot Symbols** for atoms and ions.

**Chemical bond:** what holds atoms or ions together in a compound

**Ionic bond:** electrostatic attraction holding cations and anions together in an ionic compound

**Ionic compounds** exist as 3D networks of ions.

* The formula indicates the **ratio** of ions.
* At 25°C ionic compounds are solids with very high melting points.

**Electronegativity (EN):** ability of an atom **in a bond** to draw electrons to itself

* Know F is most electronegative; further away from F, less electronegative an atom.

**Covalent bond:** sharing of electrons between two nonmetal atoms

**Nonpolar covalent bond:** equal sharing of electrons by two atoms with equal EN

* Electron density is even distributed between the two atoms.

**Polar covalent bond:** unequal sharing of electrons by 2 atoms with different EN

values → **dipole** (separation of charges)

* Electron density is concentrated towards the more electronegative atom.
* **dipole moment:** quantitative measure of the polarity of a bond.
* The greater the EN difference, the more polar the bond.

→ Rank different bonds in terms of increasing polarity.

Use **delta notation (**δ**+ and** δ**-) and a dipole**

* **Arrow** to indicate which atom in a bond is more electronegative.

**Octet rule:** atoms bond such that each has 8 electrons, except H only needs 2 electrons.

**Draw Lewis Structure for Molecules**

1. Count total number of valence electrons, then divide by 2 to get # of electron pairs.
2. Draw skeleton structure

* Least electronegative element in center, and H and F are always outer atoms

1. Connect all the atoms by drawing lines to represent single bonds

* Distribute remaining electrons around outer atoms, then central atom until each has an octet.
* Make double or triple bonds only if an atom does not have an octet.
* H and F only form single bonds.

**Resonance structures:** two or more structures representing a single molecule with delocalized electrons that cannot be

Described fully with only one Lewis structure

* Recognize which molecules require resonance structures.
  + Know “delocalized electrons” are shared by three or more atoms in a molecule.
  + Note: The delocalized electrons do NOT oscillate between atoms. They are **always** shared by the atoms, but there’s no other way for us to represent this.
  + Know the relative length and strength of bonds with delocalized electrons.

**Molecular Geometry (or Shape)**

* Use Lewis structure and to get 3D shape and bond angles:
  + AX2 → linear → 180°
  + AX3 → trigonal planar → 120°
  + AX4 → tetrahedral → 109.5°
  + AX2E → bent → <120°
  + AX3E → trigonal pyramidal → <109.5°
  + AX2E2 → bent → <109.5°

1. Breathalyzers estimate the amount of alcohol in the blood by measuring the alcohol in the breath. The breathalyzer uses the redox reaction below to determine the amount of alcohol in the blood. Answer the following questions using this balanced chemical equation.

3 C2H5OH + 10 H2CrO4 🡪 3 CH3CO2H + 8 Cr2(CrO4)3 + 13 H2O + 2674 kJ

ethanol chromic acid acetic acid chromium(III) chromate water

* 1. How many moles of chromic acid can react with 5.92 moles of ethanol?
  2. How many mg of ethanol are in a sample of a driver’s breath that produces 21.8 mg of acetic acid?
  3. How much energy will be produced if 38.1 grams of chromic acid react with excess ethanol?
  4. If 62.8 grams of ethanol react with excess chromic acid to produce 89.4 grams of water, what is the percent yield of the reaction?
  5. If 9.25 g of ethanol are allowed to react with 62.8 g or chromic acid, how many g of acetic acid should be produced?

1. Write a balanced equation for the reaction, if any that occurs in each of the following cases. Assume that all soluble reactants are added in the form of aqueous solutions. Indicate gases and precipitates that are formed, as well as insoluble solid reactants. If no reaction occurs, then write **NO RXN**, and do not write a balanced equation. Be sure to **balance** your equations and include your **phase labels**.
2. potassium hydroxide + sodium hydroxide 🡪
3. sodium acetate + hydrochloric acid 🡪
4. zinc bromide + potassium phosphate 🡪
5. hydrochloric acid + calcium 🡪
6. nitric acid + barium hydroxide 🡪

1. ammonium nitrate + sodium hydroxide 🡪
2. How many electrons are there in an orbital? \_\_\_\_\_\_\_\_\_

In an s sublevel? \_\_\_\_\_\_\_\_

A p sublevel?\_\_\_\_\_\_\_\_\_

A d sublevel?\_\_\_\_\_\_\_\_\_

1. How is the line spectra of an atom produced? (Answer this on an atomic level)
2. How do atomic orbitals fill? If a p sublevel has 4 electrons, which orbitals will they occupy? Draw the sublevel using arrows to represent electrons and show spin based on the direction of the arrow.
3. Explain why each of the following electron configurations for a p sublevel is disallowed.
4. Write the complete and shorthand electronic configuration for the following atoms and ions.

|  |  |
| --- | --- |
| shorthand | complete |
| Si: | Si: |
| Mg+2 : | Mg+2 : |
| S‑2: | S‑2: |
| V: | V: |
| Mn: | Mn: |
| Ni +2: | Ni +2: |

1. Explain why an atom might have an anomalous configuration. Predict which elements might have anomalous configurations.
2. How many valence electrons in an atom of phosphorous? Of barium? Of krypton?
3. Explain how an ionic bond differs from a covalent bond.
4. Explain how a pure covalent bond differs from a polar covalent bond.
5. What is the lewis electron dot structure of carbon? Of arsenic? Of potassium?
6. Rank the following atoms in order of increasing radius:

\_\_ N \_\_ Sb \_\_ Rb \_\_ Cs \_\_ F

1. Name the element that corresponds to each of the following:
2. alkali metal with the smallest atomic radius\_\_\_\_\_\_\_\_\_\_
3. Group 5A element with the highest ionization energy\_\_\_\_\_\_\_\_\_\_\_\_
4. [Kr] 5s24d10\_\_\_\_\_\_\_\_\_\_\_\_\_
5. Give three anions that are isoelectronic with neon?
6. List the following in increasing (L to R) ionization energy.

N, Ca, Cl, Fr, Rb

1. Arrange the following elements in order of increasing atomic size: Ca, b, S, Si, Ge, F
2. Arrange the following elements in order of increasing metallic character: Fr. Sb. In, S, Ba, Se
3. Draw Lewis Electron Dot Structures for the following molecules and tell the orbital and molecular geometries for each

|  |  |
| --- | --- |
| * 1. PH3 | * 1. C2H4 |
| * 1. CBr4 | * 1. C2H6 |
| * 1. OF2 | * 1. N2H2 |
| * 1. SCl2 | * 1. C2H2 |
| * 1. CO2 (C is central) | * 1. Cl2CO (C is central) |

1. Draw Lewis Electron Dot Structures for the following ions. Include resonance structures as necessary.

|  |  |
| --- | --- |
| * 1. CN-1 | * 1. NO2-1 |
| * 1. SO3-2 | * 1. CO3-2 |
| * 1. NH4+1 | * 1. ClO3- |

1. Predict the orbital and molecular geometry around each of the starred atoms in the drawing below:

