# NAME \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ CHEMISTRY FINAL EXAM REVIEW

# DATE \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

DIRECTIONS: Go through the review and try to answer as many questions as you can without having to look anything up. Circle/highlight the questions you couldn’t answer right away. This will identify areas that you need to concentrate on for the final. Go through your notes and worksheets to help you answer the rest of the questions.

***Chapters 2 and 3 – MATTER***

1. What’s matter?

Anything that has mass and takes up space

1. What are the two forms of matter?

Pure substances and mixtures

1. What are the two types of mixtures?

heterogeneous – non-uniform throughout

homogeneous – uniform throughout

1. What is the difference between the two types of mixtures?

see above

1. Give examples of the two types of mixtures.

hetero = mint chocolate chip ice cream

homo = chocolate ice cream

1. What are the two types of Pure Substances?

elements and compounds

1. What is the difference between the two types of Pure Substances?

element = 1 type of atom

compound = multiple

1. Give examples of the two types of Pure Substances.

iron = element  
water = compound

1. What is an atom?

the simplest type of matter than cannot be broken down by physical/chemical means

1. What is a molecule?

a piece of matter containing two or more atoms

1. What is a chemical property? Give two examples.

A property of matter than can only be observed while changing the substance – rusting, burning, reacting

1. What is a physical property? Give two examples.

A property of matter than can be observed without changing the identity of the matter – boiling, melting, mixing, painting, freezing, etc

1. What is density?

mass / volume = a measure of how many particles are packed into a space

1. What is the density of a liquid that has a mass of 50. g and a volume of 300. mL?

0.17 g/mL

1. A metal has a density of 8.92g/mL. What mass will 33.4 mL of the metal have?

298 g

1. A metal has a density of 2.7 g/mL. 50.8 g of the metal will occupy what volume?

19 mL

1. What is an intensive physical property? Give two examples.

One that’s true no matter how much “stuff” is present – density, specific heat, color, melting point

1. What is an extensive physical property? Give two examples.

One that’s dependent on how much “stuff” is present – length, mass, size, weight

1. What’s the difference between a qualitative observation and a quantitative one?

qualitative = not numerical

quantitative = numerical

1. Give examples of each.

the desk is heavy = qualitative

the desk has a mass of 28.2 kg = quantitative

1. For the letters below, write out how many significant figures they have and rank them in uncertainty from least to most:
   1. 1.00030 = 6 sig figs
   2. 0.0054 = 2 sig figs
   3. 250.0 = 4 sig figs
   4. 300 = 1 sig fig
2. Define the following properties and classify them as physical or chemical: luster, malleability, ductility, and flammability.
   1. Luster = shine = physical
   2. Malleability = ability to be pounded into a thin sheet = physical
   3. Ductility = ability to be stretched/drawn into a wire = physical
   4. Flammability = how easily something will catch fire = chemical
3. Define the law of multiple proportions and provide examples of two compounds that illustrate the concept.

The law of multiple proportions states that elements can combine in different whole number ratios to form unique compounds. An example of this law is that H2O is different than H2O2 even though they both contain the same elements.

1. Define the law of definite proportions in your own words.

The law of definite proportions states that any compound will always contain the exact composition of elements in the exact same ratio. That is, water will always be H2O.

***Chapters 4 and 5 – ATOM/SUBATOMIC PARTICLES***

1. What are the three subatomic particles that make up an atom?

proton, neutron, electron

1. Where are the three subatomic particles located in an atom?

proton and neutron = inside nucleus

electron = cloud outside nucleus

1. What are the charges of the three subatomic particles?

proton +1, neutron 0, electron = -1

1. What are the masses of the three subatomic particles?

proton = 1 amu, neutron = 1 amu, electron = 0 amu

1. What is the atomic number of an atom?

# of protons

1. What is the mass number of an atom?

# of protons + # of neutrons

1. What is an ion?

charged particle – one in which # of protons ≠ # of electrons

1. How does an atom become an ion?

gains/loses electrons

1. What are the two types of ions and how are they different?

cation = +

anion = -

1. What’s an isotope?

One type of an atom. Isotopes of the same atoms have varying #s of neutrons

1. Carbon-12 has how many protons, neutrons, and electrons?

6 PROTONS, 6 NEUTRONS, 6 ELECTRONS

1. Chlorine has two istopes: Cl-35 and Cl-37. Which is more common? How do you know?

Cl-35, THE MOLAR MASS OF CHLORINE IS CLOSER TO 35 THAN 37

1. Magnesium has three isotopes. Magnesium-24 makes up 78.99% of all magnesium atoms and has an atomic mass of 23.985. Magnesium-25 makes up 10.00% and has an atomic mass of 24.986. Magnesium-26 makes up 11.01% and has a mass of 25.982. What’s its atomic mass?

24.31 g/mol

***Chapters 5 and 6 – THE ELECTRON/PERIODIC TABLE***

1. How many energy levels are there? How many electrons can fit in each?

7 total energy levels: energy level #1 can hold 2 electrons, E level #2 can hold 8 electrons, E level #3 can hold 18 electrons, E levels #4 - #7 can all hold 32 electrons. (Some of this is theoretical. The 6th and 7th energy levels have only so far held electrons in the “s” and “p”. In theory, more sublevels than just s, p, d, and f will exist in these higher levels.)

1. How many sublevels are there? What are they? How many electrons fit in each?

4 sublevels: s can hold 2 electrons, p can hold 6 electrons, d can hold 10 electrons, and f can hold 14 electrons

1. How many orbitals are in each sublevel? How many electrons fit in one orbital?

s sublevel has only 1 orbital, p sublevel has 3 orbitals, d sublevel has 5 orbitals, f sublevel has 7. All orbitals hold only 2 electrons.

1. What does the electron configuration tell you about an atom?

It tells you how the electrons are distributed around the nucleus of an atom.

1. Color the following Periodic Table according to sublevels.

Orange/yellow = s block

Purple = p block

Green = d block

Red = f block

1. Write the electron configuration and the orbital diagrams for the following elements.
   1. Cl = 1s22s22p63s23p5
   2. Mn = 1s22s22p63s23p64s23d5
   3. Au = 1s22s22p63s23p64s23d104p65s24d105p66s24f145d9
   4. Fr = 1s22s22p63s23p64s23d104p65s24d105p66s24f145d106p67s1
   5. S = 1s22s22p63s23p4
   6. K = 1s22s22p63s23p64s1
2. What are valence electrons?

Electrons that are in the outermost energy levels

1. How many valence electrons do the following elements have?
   1. Na: 1
   2. O: 6
   3. Cu: 2
   4. P: 5
   5. Kr: 8
2. How many electrons make an atom stable? What rule tells us this?

Most elements “want” 8 valence electrons. This is called the octet rule. There are some exceptions though in hydrogen, helium, lithium, beryllium, and boron (they “want” 2 instead)

1. What type of ion forms if an atom gains electrons?

Negative, or anion

1. What type of ion forms if an atom loses electrons?

Positive, or cation

1. What happens when an atom absorbs energy?

Electrons jump up in energy levels

1. What happens with an atom then releases that energy?

The energy gets released in the form of light as the electrons jump back down in energy level.

1. Determine the charge that the following elements will take to become stable.
   1. Li: +1
   2. N: -3
   3. Si: +/-4
   4. Ca: +2
   5. F: -1
   6. Al: +3
2. Describe the trend for electronegativity on the periodic table (which elements have high electronegativities and which have low electronegativitities).

Electronegativity is a measure of how much an atom “wants” electrons. The elements in the top right of the periodic table have very high electronegativities whereas the elements at the bottom left of the table have very low electronegativities.

1. Describe the trend for atomic radius on the periodic table (which are big and which are small).

The atomic radius trend is the opposite of electronegativity. The smallest atoms are at the top right and the biggest atoms are at the bottom left.

1. Describe the trend for *ionic* radius on the periodic table.

When atoms become ionized, they tend to get smaller. They do this due to fewer electron shells (cations) or greater electrostatic attractive forces between the nucleus and the electrons (anions).

***Chapters 8 and 9 – IONIC AND MOLECULAR COMPOUNDS***

1. What is an ionic compound?

A compound that contains a metal bonded to a nonmetal (or one that involves a polyatomic ion). The bond is one in which electrons are given transferred from one atom to another.

1. What is a molecular compound?

A compound that is joined with a covalent bond. It contains only nonmetals. The bond involves electrons being shared between atoms.

1. How do you determine the formula of an ionic compound?

Balance the charges (criss-cross rule?)

1. Which ion is written in the formula first?

The cation

1. What is a polyatomic ion?

A group of bonded atoms that collectively have an overall charge

1. Name and write out the formulas for all seven polyatomic ions you need to know.

Ammonium = NH4+ Hydroxide = OH- Nitrate = NO3-

Carbonate = CO3-2 Sulfate = SO4-2 Phosphate = PO4-3

Acetate = C2H3O2-

1. Write an “I” for ionic or a “C” for covalent characteristics.

\_\_I\_\_\_high melting point \_\_I\_\_\_crystalline structure

\_\_C\_\_\_two nonmetals \_\_\_I\_\_good conductors

1. Would you expect an ionic compound or covalent compound to dissociate in water? Why? What does it mean to dissociate?

Ionic compounds dissociate in water. It means that they split up into their component ions. For example, sodium chloride, an ionic compound, breaks up into Na+ and Cl-.

1. List the covalent prefixes for numbers 1-10.

1 - mono 3 - tri 5 - penta 7 - hepta 9 – nona

2 - di 4 - tetra 6 - hexa 8 - octa 10 - deca

1. Determine the formula of the following compounds. (Remember, you must first determine if the compound is molecular or ionic!)
2. Magnesium Chloride MgCl2 m. Hydrogen Nitride H3N
3. Iron(II) Sulfide FeS n. Beryllium Chloride BeCl2
4. Barium Nitrate Ba(NO3)2 o. Manganese(II) Selenide MnSe
5. Dihydrogen Monoxide H2O p. Lead(IV) Sulfide PbS2
6. Sodium Sulfate Na2SO4 q. Tetracarbon Hexaiodide C4I6
7. Aluminum Fluoride AlF3 r. Silver Carbonate Ag2CO3
8. Lithium Acetate LiC2H3O2 s. Pentasulfur Octafluoride S5F8
9. Ammonium Bromide NH4Br t. Chromium(III) Acetate Cr(C2H3O2)3
10. Rubidium Oxide Rb2O u. Hydrogen Monobromide HBr
11. Copper(III) Phosphate CuPO4 v. Vanadium(V) Hydroxide V(OH)2
12. Calcium Chloride CaCl2 w. Gallium(III) Oxide Ga2O3
13. Nitrogen Trihydride NH3 x. Cesium Sulfate Cs2SO4
14. Write the name of the following compounds. (Remember, you must first determine if the compound is molecular or ionic.)
15. SrCl2 strontium chloride m. CdO cadmium(II) oxide
16. NiS nickel(II) sulfide n. K3PO4 potassium phosphate
17. PBr5 phosphorus pentabromide o. NaF sodium fluoride
18. Ba(NO3)2 barium nitrate p. CrSe chromium(II) selenide
19. Mg3(PO4)2 magnesium phosphate q. FeCO3 iron(II) carbonate
20. BI3 boron triiodide r. Al(NO3)3 aluminum nitrate
21. Zr(C2H3O2)2 zirconium(II) acetate s. SrO strontium oxide
22. H2S hydrosulfuric acid t. PtBr2 platinum(II) bromide
23. K2Se potassium selenide u. Be3N2 beryllium nitride
24. CaF2 calcium fluoride v. ZnI2 zinc iodide
25. NH4Cl ammonium chloride w. PbS2 lead(IV) sulfide
26. BaCO3 barium carbonate x. Cu(OH)2 copper(II) hydroxide