Summer Review for AP Chemistry

This material is intended as a review in preparation for AP Chemistry. Topics listed will be covered in the first weeks of AP Chemistry. Please review the information in this packet as a way to prepare for the beginning of the year.

Significant figures

All calculations will require knowledge of the proper use of significant figures. You will be required to recognize and report all answers with the proper number of significant figures.

Electron configuration

Knowledge of how the electrons fill orbitals and how this affects their behavior and placement on the Periodic Table.

Trends of the periodic Table

Electronegativity and Atomic Radius and how they affect reactivity.

Mole Mass Relationships

Topics include molar mass, problem solving and multi step problems.

Significant Figures

Counting Significant Figures

1) Nonzero integers. Always count as significant.

78564 **(5)**

2) Zeros. Three classes of zeros.

a) Leading zeroes, preceding all of the nonzero integers never count as significant.

0.00004765 **(4)**

b) Captive zeros fall between nonzero digits. Captive zeros are always significant.

39800087 **(8)**

c) Trailing zeros, right end of the number. Only if the number is written with a decimal point.

0.165400 **(6)** 10000 **(1)**

3) Exact numbers, obtained by counting or by definition. Non-limiting

5 cars, and 1 inch = 2.54 cm are said to be exact

Rules for Calculations

For multiplication or division, the number of significant figures in the result is the same as that in the measurement with the smallest number of significant figures.

4.56 x 1.4 = 6.384	**(6.4)**
8.315 / 298 = 0.279026846	**(0.279)**
2.3 x 3.15 = 7.245	**(7.2)**
5.18 x 0.0208 = 0.107744	**(0.107)**
12.6 x 0.53 = 6.678	**(6.7)**

For addition and subtraction, the limiting term is the one with the smallest number of decimal places

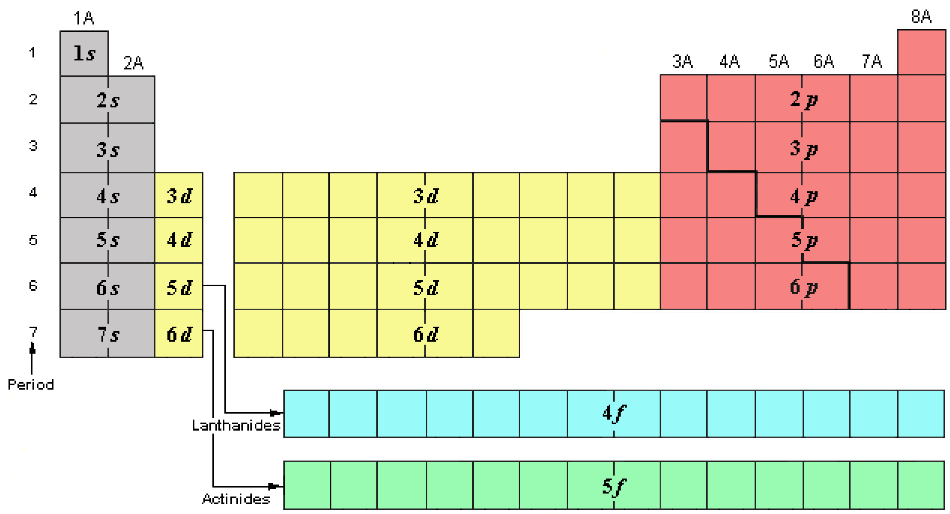
12.11 + 18.0 + 1.013 = 31.123	**(31.1)**
100 + 0.12 + 0.000086 = 100.120086	**(100)**
0.6875 – 0.1 = .5875	**(0.6)**

Rules for writing electron configurations

1) Determine the amount of electrons an element is going to have. Remember for a neutral atom the amount of electrons is the same as the atomic number (number of protons). We will learn how to write the configurations for ions later. You are going to have to account for all of the electrons when you are finished.

2) **Aufbau (build upon)** The electrons are going to go into the lowest energy levels first.

Energy Level Chart



The lowest energy level is 1s. There is only one 1s orbital.

3) Pauli Exclusion Principle. Only two electrons per orbital. We put the first electron in 1s and designate that "spin up" ↑ .

If there is another electron it will go into to 1s (remember two per orbital) and we will designate that as "spin down" which serves to differentiate it from the first electron in the orbital. ↑↓

We would write it as ↑↓
1s

4) Hund's Rule Electrons spread out before pairing up. When filling p orbitals, which have three (six electrons) and d orbitals, which have five (ten electrons) spread the electrons out before pairing up. With "p"s the first three electrons go into separate orbits spin up, the fourth, fifth and sixth electrons the go in, forming pairs, spin down. This is also the case with "d"s with the first five going into separate orbits spin up with six through ten forming pairs spin down.

Example: Writing the electron configuration for Oxygen "O"

Oxygen has eight electrons that need to be accounted for.

a) **Lowest energy level first.**

↑↓
1s

Two electrons are accounted for, four more to go.

Next energy level is 2s, refer to the energy level chart.
We place our two electrons in spin up and spin down.

↑↓ ↑↓
1s 2s

This accounts for four electrons, two more to go

Our next energy level is 2p, refer to the energy level chart. There are three "p" orbitals.

___ ___ ___
2p

These can accommodate six electrons, two per orbital.
The first electron goes into the first orbital spin up.

↑ ___ ___
2p

Now we have the first application of "**Hund's Rule**," electrons spread out before pairing up. The second electron goes into the second p orbital, spin up, third electron goes into the third p orbital, spin up. Now the fourth and final electron goes into the first p orbital spin down.

↑ ↑ ___	↑ ↑ ↑	↑↓ ↑ ↑
2p	2p	2p

At this point we have finished the electron configuration for oxygen, all eight electrons are accounted for.

↑↓ ↑↓ ↑↓ ↑ ↑
1s 2s 2p

Electron configuration for oxygen.
With the unpaired electrons we would consider this "Paramagnetic"

Trends on the Periodic Table

Electronegativity

Electronegativity is the ability of an atom of an element to attract electrons when in a compound.

Electronegativity values range from a high of 4 (Fluorine) to a low of .7 (Francium) according to the Pauling electronegativity scale.

Differences in electronegativity values are used to classify bond types as either covalent, polar covalent or ionic.

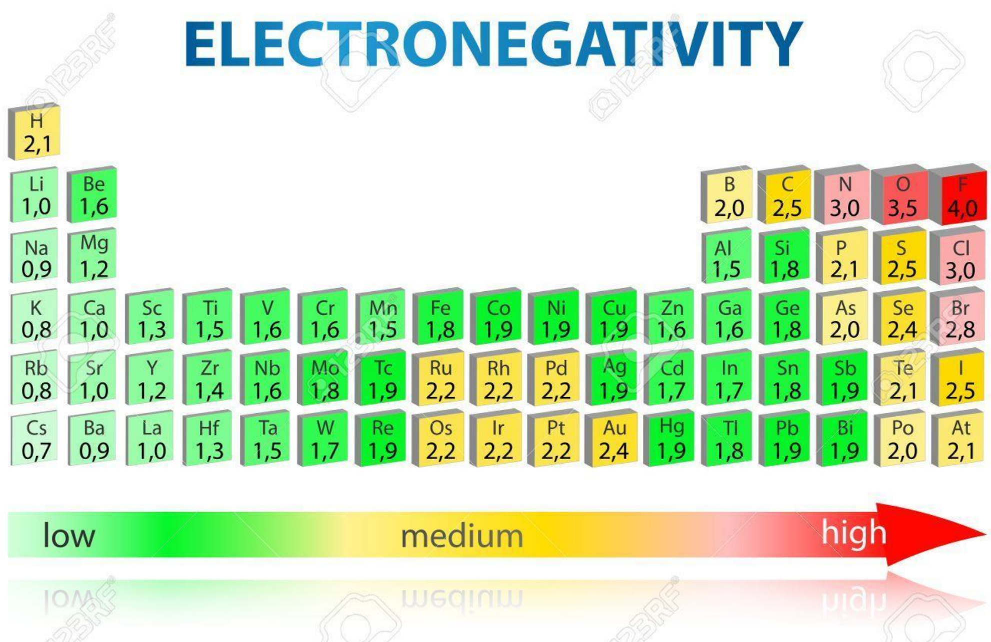
Electronegativity increases as we go from left to right across the periodic table.

Metals, elements on the left side of the periodic table, like to give away their electrons to complete their outer shell. If their tendency is to give away their electrons they generally do not want to attract electrons to themselves.

Non-metals, elements on the right side of the periodic table, like to acquire electrons to complete their outer shell. If the tendency is to acquire electrons, the elements will have a higher electronegativity value.

Electronegativity decreases as you go from top to bottom down a column.

As you go down a group you add shells which shield the outermost electrons from the attractive forces of the nucleus. This reduces the ability to attract electrons towards the nucleus.



Atomic Size

Coulomb's equation: $$F = k_e \frac{q_1 q_2}{r^2}$$

q - represents charge and the force is attractive if the charges are opposite.

r - represents distance

Atomic Radius

Radius increases as you go from top to bottom on the periodic table.

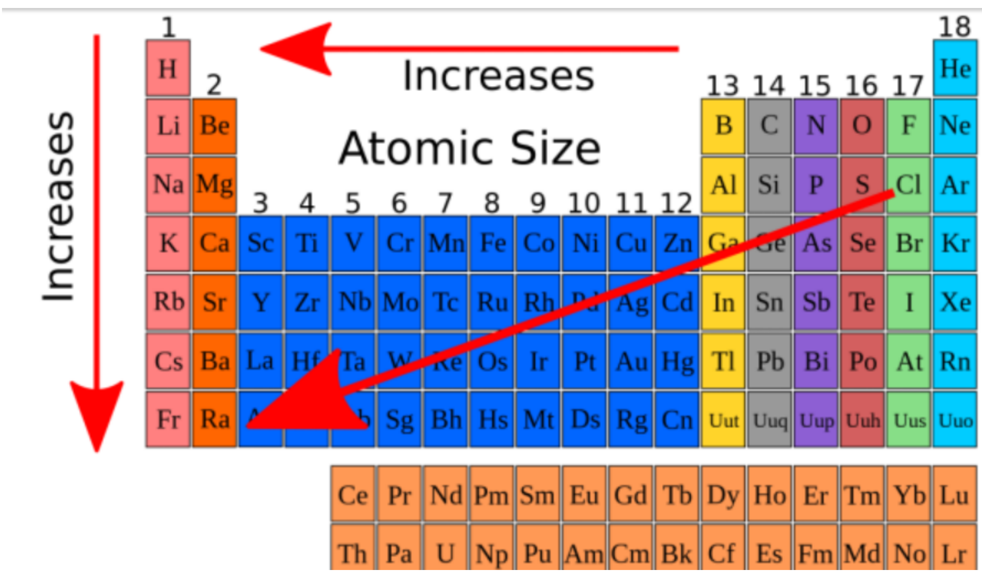
As we go down a period we add shells. Remember from filling electron orbital diagrams. Each shell added increases the distance of the outermost electrons from the nucleus.

From Coulomb's equation we are increasing r while keeping q constant. This leads to a smaller attractive force and allows for the outermost electrons to be held less tightly. This leads to a larger radius.

Radius decreases as you go from left to right on the periodic table.

As we go left to right across a row we add protons. This serves to increase effective nuclear charge without adding additional shells.

From Coulomb's equation we are increasing q (charge) while keeping r (distance) constant. This leads to a larger attractive force and allows for the outermost electrons to be held more tightly. This leads to a smaller radius.



Mole Mass Relationship

Molar Mass of a Compound

The mass of a mole of an element is its molar mass.
Calculate the molar mass of H_20. Use proper amount of significant figures.
From the periodic table we find that Hydrogen "H" has an atomic mass of 1.008.
By looking at the molecular formula of H_20 we see there are 2 hydrogens.

$2 \times 1.008 = 2.016$

Again from the molecular formula we see there is 1 oxygen and from the periodic table we see "O" has an atomic mass of 16.00
$1 \times 16.00 = 16.00$

We add the two, 2.016 + 16.00 = 18.016 or 18.02

The units are g/n 18.02 g/n or 1 mole of H_20 weighs 18.02 grams or 1 mole of H_20 = 18.02 grams

What this means is that 6.02×10^{23} molecules of H_2O weigh 18.02 grams.

Calculate the molar mass of $Mg(OH)_2$. Use proper amount of significant figures.

From the periodic table we find that Magnesium "Mg" has an atomic mass of 24.30 and from the formula we see that there is one "Mg"

1 x 24.30 = 24.30

By looking at the molecular formula of $Mg(OH)_2$ we see that (OH) has a subscript of 2. This means that there are 2 oxygens 2 x 16.00 = 32.00

And 2 Hydrogens 2 x 1.008 = 2.016

We add the three amounts, 24.30 + 32.00 + 2.016 = 58.316 or 58.32

The units are g/n 58.32 g/n or 1 mole of $Mg(OH)_2$ weighs 58.32 grams or 1 mole of $Mg(OH)_2$ = 58.32 grams

Solving Dimensional Analysis Problems

1) Identify what you are looking for, it will generally be mass in grams, volume of a gas in Liters or a number, atoms or molecules.
2) Identify your starting point. This is an amount.
3) Look for an equality statement, something equals something, that allows you to at the very least cancel a unit and at the most, find your endpoint. Examples might be g/n 1n=22.4L or 1n= 6.02×10^{23}

Example Problem: What is the mass in grams of 2.40×10^{24} molecules of CO?

We are looking for mass in grams, so grams will be the unit that you will end with.
Our starting point is 2.40×10^{24} molecules of CO

2.40×10^{24} molecules of CO x

Now we look for a conversion or equality statement that at the very least will eliminate molecules. Avagadro's number will work. Even though it does not get you to grams it eliminates molecules

1n = 6.02×10^{23} molecules

$$2.40 \times 10^{24} \text{ molecules of CO} \times \frac{1n}{6.02 \times 10^{23} \text{ molecules}} \times$$

1n goes in the numerator and 6.02×10^{23} goes into the denominator. This allows you to cancel molecules.
Now we need to find a conversion factor that will allow us to go from moles to grams.

Molar mass. Calculate the molar mass of CO. C =12.01 g/n and O=16.00 g/n
The molar mass of CO written to the correct amount of significant figures is CO= 28.01 g/n

$$2.40 \times 10^{24} \text{ molecules of CO} \times \frac{1n}{6.02 \times 10^{23} \text{ molecules}} \times \frac{28.01g}{1n} = 112 \text{ grams CO}$$

We place moles in the denominator to cancel moles and 28.01grams in the numerator. The answer with the correct amount of significant figures is 112 grams.