Chemistry I - Standard

Atomic Theory & Nuclear Chemistry Notes

The idea that all of matter – any substance which contains mass – is composed of small, microscopic entities known as “atoms” was not always thought to be a universally held truth. Like most competing theories of our day – the cause of cancer, the human-directed proliferation of global warming, or even the thought of life elsewhere in the universe – scientists debated what constituted matter. By now, we have an incredibly specific picture of what we *think* matter is composed. This definition, while consistent in its foundation, is ever-changing on the fringe.

**I. Development of the Modern Atomic Theory**

* Matter is defined as: \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_
* All of matter is composed of small, microscopic entities known as atoms. The word atom comes from the Greek work “atomos” which means \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_. This word was coined by the philosopher Democritus, who thought that atoms were the smallest unit of matter. We would find out later that he was wrong. Atoms are composed of three subatomic particles: (1) \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_, (2) \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_, and (3) \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_.
* Table of Subatomic Particle Properties

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
| *Particle* | *Representation* | *Charge* | *Location* | *Relative Mass* |
| proton |  |  |  |  |  |
| neutron |  |  |  |  |  |
| electron |  |  |  |  |  |

* The Modern Atomic Theory states that the atom is an electrically \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_, spherical entity composed of a positively charged \_\_\_\_\_\_\_\_\_\_\_\_\_ \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ surrounded by one or more negatively \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ electrons.

**II. Atomic Symbol Representations**

In our study of the properties of different atoms, it will be increasingly important to represent these atoms quickly and concisely. We do not want to encumber ourselves with too many words. In this light, scientists have developed a system for representing these atoms in an efficient fashion:

X = atomic symbol of the element

$\begin{matrix}A\\Z\end{matrix} $X

Z = atomic number

A = mass number; A = Z + N

N = number of neutrons in the nucleus

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* Determining the number of protons, neutrons, and electrons:

a) protons: \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

b) neutrons: \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

c) electrons: \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ (neutral atom)

Example Set #1: Finding the number of protons, neutrons, and electrons:

|  |  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- | --- |
|  | Nuclear Name | Nuclear Symbol | Atomic Number | Mass Number | #Protons | # Neutrons | # Electrons |
| a | Hydrogen – 1  |  | 1 |  | 1 |  | 1 |
| b |  | $\begin{matrix}15\\7\end{matrix}$N |  |  | 7 | 8 |  |
| c | Aluminum – 27  |  |  | 27 |  |  | 13 |

* In certain instance, atoms may have a different mass number. When an element has a mass number that differs from the one found from the periodic table, it is known as an isotope of that element.
* An isotope is when atoms of the \_\_\_\_\_\_\_\_\_\_ element have the same number of \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_, but a different number of \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_.

Example Set #2: Isotopes

|  |  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- | --- |
|  | Nuclear Name | Nuclear Symbol | Atomic Number | Mass Number | #Protons | # Neutrons | # Electrons |
| a | Hydrogen – 1  |  | 1 |  | 1 |  | 1 |
| b | Hydrogen – 2  |  |  |  |  |  |  |
| c | Hydrogen – 3  |  |  |  |  |  |  |

* Remember: the only value for an atom that absolutely has to been the same is the **atomic number**. The other values may change and this is perfectly acceptable. A common mistake most first-time chemistry students make is they assume the mass number will **always** be the same (and come from the periodic table). Always look at the nuclear name (or A + N) to get the mass of the atom.

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So far in our study of matter, we’ve discussed how the proton number never changes for a specific atom (if it did it would be a completely different element), how the number of neutrons changes (these are called **isotopes**), but how do we indicate when the number of ­ELECTRONSchanges?

When the number of electrons changes from the original, electrically neutral number these types of atoms are called \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_.

* Ions are either positively or negatively charged.
* Positively charged ions are called \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ and means the atom has \_\_\_\_\_\_\_\_ electrons.
* Negatively charged ions are called \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ and mean the atom has \_\_\_\_\_\_\_\_\_ electrons.

$\begin{matrix}A\\Z\end{matrix} $X #±

* We indicate the charge of the atom by placing the number in the upper right-hand corner of the nuclear symbol.
* We indicate the charge in the isotope name by placing the number and sign (either positive or negative) in parentheses next to the mass number. Let’s look at some examples:

Example Set #3: Ions

|  |  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- | --- |
|  | Nuclear Name | Nuclear Symbol | Atomic Number | Mass Number | #Protons | # Neutrons | # Electrons |
| a | oxygen – 16 (anion, 2-) |  |  |  |  |  |  |
| b | zinc – 65 (cation, 2+) |  |  |  |  |  |  |
| c | lead – 214 (cation, 4+) |  |  |  |  |  |  |
| d |  |  |  | 86 | 33 |  | 36 |

* The notation of ions may seem confusing. If you are asking yourself, why the number of electrons ***increases*** when the charge of the atom is a ***negative*** number, then you are asking yourself the right question. Many scientists (myself included!) have asked this same question. This idea seems counter-intuitive. Let’s see why by drawing atomic diagrams below. Let’s use Li-7 and Li-7 (cation, 1+):

|  |
| --- |
|  |

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**III. Average Atomic Mass**

By now, we know that all of matter is composed of \_\_\_\_\_\_\_\_\_\_\_\_\_ and that these entities are made up of the three \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ particles. If the number of protons and electrons are \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_, then we classify this is an electrically \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ atom. If not, then we classify this substance as an \_\_\_\_\_\_\_\_. When the mass number of the atom does not match the mass on the periodic table, we say this is an \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ of that atom. While these atoms have the same atomic number, they differ in their number of \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_.

* The number located on the bottom of each box on the periodic table is an **average** of all of the masses of the known isotopes of that element.
* In order to calculate the average atomic mass of an element, scientists need to know three things:
1. The number of isotopes for that element
2. The mass of each isotope.
3. The percentage abundance of each isotope.
* The percentage abundance is a measurement of the proportions in which the isotopes exist.
* Not all isotopes of an element have the same percentage abundance. Take hydrogen for example. Hydrogen has three naturally occurring isotopes. (pictured above) These three isotopes exist in different amounts. Notice that the hydrogen-1 has the highest percentage abundance. This is the most common isotope of the element.
* To calculate the average atomic mass for any element, following this equation:

Average atomic mass = (Isotopic mass x percent abundance)1 +

 (Isotopic mass x percent abundance)2 +

 (Isotopic mass x percent abundance)3 +

 (Isotopic mass x percent abundance)n where n = the number of isotopes.

Examples:

1. Hydrogen-1 occurs in 96% of a sample, Hydrogen-2 occurs in 3% and Hydrogen-3 in 1%. Calculate the average atomic mass for hydrogen.

(1\*0.96) + (2\*0.03) + (3\*0.01) = **1.05 amu**

1. The element copper has naturally occurring isotopes with mass numbers of 63 and 65. The relative abundance and atomic masses are 69.2% for a mass of 62.93amu and 30.8% for a mass of 64.93amu. Calculate the average atomic mass of copper.

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Examples: (continued)

1. Calculate the average atomic mass of sulfur if 95.00% of all sulfur atoms have a mass of 31.972 amu, 0.76% has a mass of 32.971amu and 4.22% have a mass of 33.967amu.
2. Calculate the average atomic mass of bromine. One isotope of bromine has an atomic mass of 78.92amu and a relative abundance of 50.69%. The other major isotope of bromine has an atomic mass of 80.92amu and a relative abundance of 49.31%.

**IV. Introduction to Nuclear Chemistry**

* **Radioactivity:** the release of nuclear radiation in the form of particles & rays from a radioactive element.
* Isotopes are often unstable – they have more neutrons than the element “wants”
* The isotopes are naturally occurring & decompose at different rates depending on the type of element – no outside influence is needed to make this decay happen.
* A nucleus that is unstable can become stable by undergoing nuclear decay
* Properties: Alters photographic film; produce fluorescence in some compounds, electric charge can be detected in the air surrounding radioactive elements; damages cells in most organisms.
* As the atoms decompose, they give off alpha, beta & gamma particles
* The **Law of Conservation of Matter** states that “Matter cannot be created nor destroyed under normal chemical or physical means.” That said, the number of protons, neutrons & electrons in the beginning have to equal the number of protons, neutrons, and electrons in the end.
* Types of Radioactive Particles:
1. **Alpha Particles (**α **-ray)**
	* + 2 protons + 2 neutrons. (42He)
		+ Remaining atom has an atomic number 2 less than the original atom & an atomic mass 4 less than the original atom.
		+ Weakest type of radiation – can be stopped by a sheet of paper.
		+ $\begin{matrix}235\\92\end{matrix}$ U → $\begin{matrix}231\\90\end{matrix}$ Th + $\begin{matrix}4\\2\end{matrix}$ He
2. **Beta Particles (β-ray)**
	* + High-speed “electron” that is formed inside the nucleus when a neutron breaks apart – can be shown as $\begin{matrix}0\\-1\end{matrix}$ β or $\begin{matrix}0\\-1\end{matrix}$ e
		+ Can pass through 3 mm of aluminum; strong, but not lethal
		+ $\begin{matrix}231\\90\end{matrix}$ Th → $\begin{matrix}231\\91\end{matrix}$ Pa + $\begin{matrix}0\\-1\end{matrix}$ e

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* Types of Radioactive Particles: (continued)
1. **Gamma Rays (γ-ray)**
	* + Energy that is given off.
		+ The most penetrating & dangerous of the radiation – can pass through several cm of lead
		+ Does not change the atomic number or mass of an element – it has no mass at all, it’s just energy.
		+ $\begin{matrix}235\\92\end{matrix}$ U → $\begin{matrix}231\\90\end{matrix}$ Th + $\begin{matrix}4\\2\end{matrix}$ He + 2$\begin{matrix}0\\0\end{matrix}$ γ

**V. Types of Nuclear Reactions**

1. In a ***fission*** reaction, a bombarding particle splits a nuclide into two nearly equal nuclides. Often, one or more expelled particles are also emitted. These reactions are highly exothermic.
* Gives off a huge amount of energy.
* For fission reactions to be sustained there must be a “critical mass” present – enough of the radioactive material present so that a chain reaction takes place.
* **Bombarding particle**: the particle (see types of particles above) used to split a large nucleus.
* Chain reactions usually occur: when the pieces “break away”, they hit new atoms and break those apart.
* Common in nuclear power plants and weaponry:
1. n + U-235 🡪 Te-137 + Zr-97 + 2 n
2. n + U-235 🡪 Ba-142 + Kr-91 + 3 n
3. In a ***fusion*** reaction, light nuclei are thrown together at extremely high speeds – high enough to cause them to “fuse” together. Without a doubt, the most common fusion reactors are the stars.
4. ***Spontaneous decay*** occurs without an outside influence. Because the nuclide is not stable, it will emit a particle on its own (presumably to try to achieve a balance of neutrons & protons)

4. The first ***transmutation*** reaction was performed by Ernest Rutherford in 1919.

1. He converted N-14 into O-17 by bombarding the nitrogen with an alpha particle, and a proton was the expelled particle.
* Balancing nuclear reactions
* Nuclear reactions show radioactive processes. They exhibit the types of nuclear reactions explained above.
* REMEMBER: The Law of Conservation of Mass!! The mass on the left side of the arrow MUST equal the mass on the right side of the arrow!

Examples:

1. $\begin{matrix}218\\84\end{matrix}$ Po → $\begin{matrix}4\\2\end{matrix}$ He + \_\_\_\_\_\_\_\_
2. $\begin{matrix}222\\86\end{matrix}$ Rn → $\begin{matrix}4\\2\end{matrix}$ He + \_\_\_\_\_\_\_\_

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1. $\begin{matrix}27\\13\end{matrix}$ Al + $\begin{matrix}4\\2\end{matrix}$ He → $\begin{matrix}30\\15\end{matrix}$ P + \_\_\_\_\_\_\_\_
2. $\begin{matrix}2\\1\end{matrix}$ H + $\begin{matrix}0\\0\end{matrix}$ γ → \_\_\_\_\_\_\_\_ + $\begin{matrix}1\\0\end{matrix}$ n
3. $\begin{matrix}239\\94\end{matrix}$ Pu + \_\_\_\_\_\_\_\_ → $\begin{matrix}1\\0\end{matrix}$ n + $\begin{matrix}242\\96\end{matrix}$ Cm
4. $\begin{matrix}6\\3\end{matrix}$ Li + $\begin{matrix}1\\1\end{matrix}$ H → \_\_\_\_\_\_\_\_ + $\begin{matrix}7\\4\end{matrix}$ Be
5. $\begin{matrix}14\\7\end{matrix}$ N + \_\_\_\_\_\_\_\_ → $\begin{matrix}1\\1\end{matrix}$ H + $\begin{matrix}14\\6\end{matrix}$ C
6. $\begin{matrix}10\\5\end{matrix}$ B + $\begin{matrix}4\\2\end{matrix}$ He → $\begin{matrix}1\\0\end{matrix}$ n + \_\_\_\_\_\_\_\_
7. In the space provided write shorthand nuclear equations for Examples 3 and 5, 6, 7, and 8. Use the notes below.
* Shorthand equations:
* Only used to show a bombarding particle hitting a radioactive isotope which results in a daughter nuclei & an emitted particle (α, β, γ, 1n, 1H)

 $\begin{matrix}129\\53\end{matrix}$ I + $\begin{matrix}1\\0\end{matrix}$ n → $\begin{matrix}0\\-1\end{matrix}$ β + $\begin{matrix}130\\54\end{matrix}$ Xe

Radioactive bombarding emitted daughter

 Isotope particle particle nucleus

* To write this in shorthand, use parentheses….

**I-129 ( n , β) Xe-130**

where the symbols & #s represent the isotope’s mass

the ( ) represent the + signs

the comma represents the →

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**VI. Half-Life**

* + Definition: The amount of time it takes for an isotope to have half of the original amount.
	+ Every radioactive isotope has it’s own half-life
	+ In order to figure out how many half lives have occurred, how long it takes for a sample to decay, the percentage remaining, the amount remaining, etc, do the following:
	1. Write down what you start with (i.e. the original amount at time = 0 (time = 0 means no decay has occurred). The number of half-lives is also 0 because no decay has occurred.
	2. After one half-life, there will be half as much (i.e. grams) as the original sample.
	3. Continue to do this until you have reached your desired amount.

\*\*Caution: don’t forget that the original amount = 0 time lapsed.

* + It is not possible to reach 0 grams – the amount just gets so small, we quit counting.

Examples and Practice:

1. The half-life of Radium-226 is 1600 years. How much of a 20 gram sample will be left after 6400 years? How many half-lives will have occurred?

|  |  |  |
| --- | --- | --- |
| Sample Amount(grams) | Half-Life | Time Elapsed(years) |
| 20 | 0 | 0 |
| 10 | 1 | 1600 |
| 5 | 2 | 3200 |
| 2.5 | 3 | 4800 |
| 1.25 | 4 | 6400 |

**Answer: After 6400 years 4 half-lives have past and there is 1.25 grams of the sample left.**

1. The half-life of radium-222 is 38 seconds. How many grams of radium-222 remain in a 12-gram sample after 76 seconds? After 114 seconds?
2. If gallium-68 has a half-life of 68.3 minutes, how much of a 10.0 mg sample is left after one half life? Two Half-lives? Three half-lives?
3. If the passing of 5 half-lives leaves 25.0 mg of strontium-90 sample, how much was present in the beginning?