# Nuclear Chemistry and Electrochemistry

*Everything You Need to Know*

This packet will discuss the main things you need to know about nuclear chemistry. After you have gone through this packet, you will be pros who can build a nuclear reactor, clean up radioactive waste, and make an atomic bomb in your garages. There will be a quiz on this stuff next Thursday!

Let's get started with nuclear chemistry!

## Part 1: What is nuclear chemistry?

Nuclear chemistry is the study of processes in which the nuclei of various atoms change in some way. There are several different ways in which this can happen:

- *Radioactive decay* is when an unstable nucleus breaks apart into smaller nuclei, or changes in some other way to make it more stable. All elements have radioactive isotopes in which the ratio of protons to neutrons makes the nuclei unstable. When radioactive decay occurs, these nuclei change into forms that are more stable.
    - All elements past uranium have NO stable isotopes. As a result, all of these elements undergo radioactive decay of some sort.
- *Nuclear fission* is a process by which a large atomic nucleus breaks up to form smaller ones. This process is accompanied by the production of a great deal of energy. This process is currently utilized in nuclear reactors and atomic bombs.
- *Nuclear fusion* is a process by which small nuclei combine to form larger ones. This process is accompanied by an even greater production of energy than nuclear fission. Nuclear fusion is the process that occurs in stars and in thermonuclear weapons (also known as "hydrogen bombs").

To review, the symbols that are used to describe atomic nuclei have the following form: ${}^{238}_{92}U$ . In this symbol (which is also described as "uranium-238" or $^{238}$U, the isotope described is an atom of uranium that has an atomic mass of 238 amu and an atomic number of 92. Because atoms of uranium *always* have an atomic number of 92, this term is usually left off of this term.

Furthermore, to find the number of neutrons in this isotope, subtract the atomic number from the atomic mass, as we discussed earlier in the year. For example, there are 146 neutrons in U-238.

<u>Sample Problems:</u>

1) How many protons, neutrons, and electrons are in an atom of U-235?
2) Describe how radioactive decay differs from nuclear fission.
3) Which process, fusion, fission, or radioactive decay produces more energy? Explain.

 

## Part 2: Nuclear decay

As mentioned before, one of the ways that nuclear reactions occur is radioactive decay. When radioactive decay occurs, the nuclei of an element either gain or lose pieces in order to gain a more stable ratio of protons to neutrons. Here are the following types of nuclear decay you need to know:

- *Alpha decay* (α): This is when a nucleus loses an alpha particle (a helium nucleus) to become more stable. One example of this reaction is the alpha decay of U-238:

  $$^{238}_{92}U \rightarrow {}^{4}_{2}He + {}^{234}_{90}Th$$

  Writing the equations of alpha decay is easy, once you know what to do. An example: Write the equation for the alpha decay of $^{244}$Pu:
    - Step 1: Write the full symbol of the nuclide that's decaying. In this case, it's $^{244}_{94}Pu$.
    - Step 2: One of the products of this reaction is an alpha particle, $^{4}_{2}He$. This gives us an equation that looks like this:

      $$^{244}_{94}Pu \rightarrow {}^{4}_{2}He + (something)$$

    - Step 3: Figure out the other product. Because the law of conservation of mass does a pretty good job of describing how the world works, the mass of the particle will be 240 (the 244 from Pu minus the 4 from the alpha particle) and the atomic number will be 92 (the 94 from Pu minus the 2 from He – this element is uranium). This gives us a final equation of:

      $$^{244}_{94}Pu \rightarrow {}^{4}_{2}He + {}^{240}_{92}U$$

- *Beta decay* (β): This is when a nucleus loses a beta particle, which is nothing more than an electron. For purposes of nuclear equations, beta particles are written with the following symbol: $^{0}_{-1}e$. An example of beta decay occurs in lead-206:

  $$^{206}_{82}Pb \rightarrow {}^{206}_{83}Bi + {}^{0}_{-1}e$$

  If you didn't know how to come up with this, simply use the same steps above to come up with the correct symbols.

- *Electron capture*: This is when a nucleus picks up an electron from the inner orbitals of the atom. Electron capture in boron-9 can be written via the following equation:

  $$^{9}_{5}B + {}^{0}_{-1}e \rightarrow {}^{9}_{4}Be$$

- *Positron emission*: This is when a nucleus emits a positron ($^{0}_{+1}e$), lowering the atomic number by one but not changing the mass. This occurs in lithium-5:

  $$^{5}_{3}Li \rightarrow {}^{0}_{+1}e + {}^{5}_{2}He$$

 

- *Gamma emission*: This is when a nucleus gives off an energetic photon of light called a gamma ray ($\gamma$). When a gamma ray is emitted by a nucleus, neither the mass nor the symbol of the element changes –however, the resulting nucleus has less energy.

Sample problems:

1) Write the equation for the nuclear reaction that occurs when copper-62 emits an alpha particle.
2) Write the equation for the reaction that occurs when americium-243 undergoes positron emission.
3) Sometimes, several nuclear reactions take place one after the other until a stable nucleus is formed. If protactinium-235 undergoes beta decay, followed by electron capture, another beta decay, gamma emission, and alpha decay, what would the resulting element be?

**Part 3: Half lives**

As has been mentioned above, many nuclei aren't very stable and undergo nuclear decay. However, if we have a single particle of a radioactive nuclide, we can't tell exactly when it will decay. We can, though, use statistics to determine roughly what fraction of particles will decay in a given length of time.

Think of it this way: Let's say that you have 100 pennies, all of which are heads. If you flip them all, you can predict that half of them will turn tails. You can't tell *which* of these will be tails, only that about half of them will be. The same things happens with nuclear decay – you can't tell which specific nuclei will decay, but you can use statistics to determine how many will decay in any period of time.

The idea that we use to describe the amount of time it takes for things to decay is that of half-life. The half life (given the symbol $t_{1/2}$) is defined as the amount of time it will take half of the radioactive particles to undergo decay. For example, if the half-life of an element is 5 minutes, and we have 80 grams of it in a container, we would expect there to be 40 grams of it after five minutes, 20 grams of it after 10 minutes, 10 grams of it after 15 minutes, 5 grams after 20 minutes, and so forth. For this compound, half of the stuff that remains goes away every five minutes.

There is, of course, an ugly equation that can be used to determine this. However, you'll find that the SOL uses examples that are simple enough that you can just figure them out using common sense. Kind of like the examples on the next page:

 

Sample problems:

1) If I have 344 grams of $^{298}Tx$ and the half-life of this element is 45 minutes, how many grams of it will remain after 270 minutes?
2) If I start with 825 grams of $^{381}Bp$ and 103 grams remain after 68 minutes, what is the half life of $^{381}Bp$?
3) The half-life of $^{421}Hi$ is 34 ms. If there are 0.00129 grams of it remaining after 544 ms, how much did I start with?

Now that you're a pro at nuclear chemistry, let's talk about electrochemistry!

**Part 4: What is electrochemistry?**

Electrochemistry refers to any reaction where any of the elements gain or lose electrons. Of course, before we can determine whether this is happening, we need to know how to figure out whether something has gained or lose electrons.

To do this, we need to determine the oxidation states of the elements in various chemical compounds. The oxidation state is really nothing more than a charge that we assign to each element, based on what it's bonded to. For ionic compounds this is simple: The oxidation state is the same as the charge it gains when it gains or loses electrons to be like the nearest noble gas. For example, in $CaBr_2$, the oxidation state of calcium is +2 and the oxidation state of each bromine atom is -1.

For other compounds, use the following rules:

- Elements that are by themselves (Fe, $Br_2$, Ca, $P_4$, and so forth) are always defined as having an oxidation state of 0.
- For covalent compounds, the overall oxidation states of all of the atoms in the molecule add up to zero when you put them together. This makes it possible to use the process of elimination to determine the oxidation states of the elements.
  - o Let's use the example of $NF_3$. The sum of the oxidation states of all the atoms in this molecule will be zero. Because fluorine is more electronegative than nitrogen, we'll arbitrarily say that it has a -1 charge (to be like the nearest noble gas). To give the molecule an overall neutral charge, nitrogen has to have an oxidation state of +3.
  - o Likewise, for $CO_2$, oxygen is more electronegative than carbon. Because oxygen has a -2 charge when it becomes like the nearest noble gas, we will assume that carbon has a +4 charge to cancel out the -2 charge of each of the two oxygen atoms.

 

- For polyatomic ions, the same theory applies as for covalent compounds, except that the overall charge needs to add up to the charge of the polyatomic ion.
  - For example, let's discuss the sulfite ion, $SO_3^{-2}$. Because oxygen is more electronegative than sulfur, we assume that each oxygen has an oxidation state of -2. Because the overall charge of the ion is -2, the charge on sulfur needs to be +4 (because +4 + (-2 x 3) = -2).

<u>Sample problems:</u> What are the oxidation states of the elements in each of the following compounds?

1) copper (II) chloride
2) $NBr_3$
3) $NO_2$
4) $SO_4^{-2}$
5) $S_8$
6) $NO_3^{-1}$
7) $NO_2^{-1}$
8) $PF_5$

**Part 5: Redox reactions**

In many chemical reactions, the oxidation state of some of the compounds changes between the products and the reactants. For example, in the reaction:

$$C + O_2 \rightarrow CO_2$$

carbon and oxygen both start the reaction with a 0 oxidation state, and in the products carbon has a +4 and oxygen has a -2 oxidation state.

*Oxidation* is the process by which a compound loses electrons. For example, because carbon went from having a 0 to +4 oxidation state, it was oxidized.

*Reduction* is the process by which a compound gains electrons. For example, because oxygen went from having a 0 to a -2 oxidation state, it was reduced.

The big question: When oxygen gained two electrons, where did they come from?

If you answered "from carbon", you're right. For this reason, in a reaction of this sort, the thing that is oxidized is called a *reducing agent* because the electrons it loses causes something else to be reduced. Likewise, if a compound is reduced, it is called an *oxidizing agent* because it grabbed those electrons from another element, causing it to be reduced.

Likewise, because oxidation cannot occur without reduction also occurring, these reactions are called *oxidation-reduction reactions*, or more commonly, *redox reactions.*

Sample problems: Some of the following reactions are redox reactions while others are not. For those reactions that are redox reactions, identify each element as having been oxidized, reduced, or neither.

1) $Fe + 2\ AgNO_3 \rightarrow 2\ Ag + Fe(NO_3)_2$
2) $Pb(OH)_2 + 2\ NaF \rightarrow PbF_2 + 2\ NaOH$
3) $CH_4 + 2\ O_2 \rightarrow CO_2 + H_2O$
4) $S_8 + 12\ O_2 \rightarrow 8\ SO_3$
5) $2\ KOH + H_2SO_4 \rightarrow K_2SO_4 + 2\ H_2O$
6) $2\ NaN_3 \rightarrow 2\ Na + 3\ N_2$
7) $XeF_4 \rightarrow Xe + 2\ F_2$
8) Based on what you know, what types of chemical reaction are most likely to be redox reactions? What types of reaction are least likely to be redox reactions?