**Learning Objectives:**

1. Define *oxidation*and *reduction*in the context of chemical reactions.
2. Assign oxidation numbers to atoms in inorganic compounds (especially binary ones).
3. Identify species that are oxidized and reduced in a redox reaction, and split the overall reaction into half-reactions.
4. Balance simple redox reactions.
5. Predict the formula of a binary salt formed from its constituent elements.

**Part I: Common Charges**

1. Recall the ionic compound NaCl. What charges do the two ions have? List their electron configurations. Why do Na and Cl “prefer” these charges?
2. What is the formula of the (ionic) compound calcium sulfide? What are the charges on the ions?
3. Binary ionic compounds are those containing a *monatomic* cation and a *monatomic* anion. The charges on these ions are also known as their “oxidation numbers.” What are the oxidation numbers of the elements in each of the following binary ionic compounds?

a) AlF3  b) TiO2 c) Cs3N

4. Given the names of the following compounds, provide their empirical formulas and the oxidation numbers of all elements present in each.

 a) Molybdenum(VI) selenide: \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

 b) Platinum(II) carbide: \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

 c) Barium iodide: \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

**Part II: Balancing Simple Redox Reactions**

1. The two compounds from Part I.1 and I.2, sodium chloride and calcium sulfide, can – theoretically – be formed from individual atoms of their constituent elements.
2. Write the chemical equation for the formation of sodium chloride from Na and Cl atoms. How many electrons are transferred?
3. Write the chemical equation for the formation of calcium sulfide from Ca and S atoms. How many electrons are transferred?
4. In nature, chlorine and sulfur do not exist as individual atoms; instead, they exist as *covalent* *molecules*, groups of atoms that are bonded together.
5. Write the *balanced* chemical equation for the formation of sodium chloride from Na atoms and Cl2 molecules. Keep in mind that the Law of Conservation of Mass requires that the number of atoms of each element is balanced – that there are the same number of Na and Cl atoms on each side of the reaction arrow. How many *total* electrons are transferred?
6. Write the *balanced* chemical equation for the formation of calcium sulfide from Ca and S8 molecules. How many *total* electrons are transferred?
7. All of the reactions you have illustrated above are *oxidation/reduction* reactions – that is, reactions that involve the **oxidation** of one species and the **reduction** of another. A substance is **oxidized** if it *loses* electrons; a substance is **reduced** if it *gains* electrons. These are also called *electron-transfer reactions* or *redox reactions*.

There are two common mnemonic devices to help you remember these definitions:

 **OIL RIG**: “**O**xidation **I**s **L**oss of electrons, **R**eduction **I**s **G**ain of electrons”

 **LEO** the lion says **GER:** “**L**oss of **E**lectrons is **O**xidation; **G**ain of **E**lectrons is **R**eduction.”

1. In the formation of NaCl, which species are *oxidized* and which are *reduced*?
2. In the formation of calcium sulfide, which species are *oxidized* and which are *reduced*?

**Part III: Half-Reactions**

1. In redox reactions, one species is always oxidized and one species is always reduced. The following is an example of a balanced redox reaction.

2 Fe + 3 Br2  🡪 2 FeBr3

 a) What are the oxidation numbers of the Fe and Br ions in the product?

 b) We assign oxidation numbers of **zero** to *pure elements* in their most stable state. So the oxidation number of metallic, elemental iron is 0, and the oxidation number of each atom of bromine in elemental bromine, Br2, is also zero. This allows us to determine the *change in oxidation number* of each element in a redox reaction. An element whose oxidation number **increases** is *oxidized*; an element who oxidation number **decreases** is *reduced*. Which element is oxidized and which is reduced in the above reaction?

c) We can “divide” a balanced redox reaction into two “half-reactions” that illustrate *only* the species being oxidized (the “oxidation half-reaction”) and the species being reduced (the “reduction half-reaction.”). These two half-reactions for this overall reaction are:

 Oxidation Half-Reaction: 2 Fe 🡪 2 Fe3+

and

Reduction Half-Reaction: 3 Br2 🡪 6 Br –

 d) You may notice that these two individual half-reactions do not obey the Law of Conservation of Charge, which states that charge is never created nor destroyed. In order to fix this, we balance each half-reaction with *electrons*. Balance the two reactions above by placing the appropriate number of electrons on either the right or left side of the reaction arrows.

**Part IV. Balancing More Complicated Redox Reactions**

Half-reactions become quite useful in balancing redox reactions. Let’s examine how using an example. The following is an *unbalanced* reaction because it does not obey the Law of Conservation of Charge:

 Cr + Pb2+ 🡪 Cr3+ + Pb

 a) Assign oxidation numbers to all the atoms and ions above.

 b) Write the individual oxidation and reduction half-reactions in the space below, making sure to balance each with the appropriate number of electrons.

Oxidation:

Reduction:

 c) In order to balance the overall equation, the total number of electrons in both half- reactions must be **the same**. Multiply each equation by some integer to make this true, then add the half-reactions back together. Write the overall *balanced* chemical equation below: