Unit 4 Name \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

Types of Reactions & Date \_\_\_\_\_\_\_\_\_\_\_\_ Block \_\_\_\_

Solution Stoichiometry

Unit 4C – Oxidation/Reduction

### Skills:

1. write & balance simple REDOX reactions
2. write & balance REDOX reactions in which H+ or OH− from solution participates in the reaction

### Notes:

Oxidation-Reduction reaction:

Originally, oxidation meant that an atom was combined with oxygen, and was therefore “oxidized”. For example:

2 Cu + O2 → 2 CuO

If we split this reaction into two “\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_,” we would have:

2 Cu0 → 2 Cu+2 + 2 e−

O20 + 2 e− → 2 O−2

In the first half-reaction, copper lost electrons to become \_\_\_\_\_\_\_\_. In the other half reaction, oxygen gained 2 electrons to become \_\_\_\_\_\_\_\_.

 oxidation:

* Which element above was oxidized; Cu or O?

 reduction:

* Which element above was oxidized; Cu or O?

Stupid Mnemonics: There is a popular mnemonic for remembering oxidation and reduction.

 *LEO the lion says ‘GER’  :* LEO stands for “\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ is Oxidation” and GER stands for “\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ is Reduction”

In a redox reaction, at least one element is oxidized, and at least one element is reduced. *An element cannot be oxidized in a chemical reaction unless some other element is reduced, and vice-versa.* (After all, the electrons have to come from somewhere, and they have to go somewhere.)

This means all \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ reactions are RedOx reactions, where something in its elemental state is trying to take on a preferred ionic state.

 oxidizing agent (or “oxidant”): the compound or ion that causes something else to be oxidized. (The oxidizing agent itself gets reduced in the reaction.)

 reducing agent (or “reductant”): the compound or ion that causes something else to be reduced. (The reducing agent itself gets oxidized in the reaction.)

 oxidation number:

**Assigning Oxidation Numbers**

* The oxidation number of a pure, neutral element is 0. (Even if it’s diatomic.)
* The oxidation numbers in a neutral compound add up to 0.
* The oxidation numbers in a polyatomic ion add up to the charge of the ion.
* The oxidation number of an ion is its charge.
* In a compound or polyatomic ion:
	+ The most electronegative element (the last one in the formula) has a negative oxidation number that is equal to the number of electrons it would need to fill its valence shell.
	+ All other atoms have positive oxidation numbers.
	+ Fluorine is always −1.
	+ Oxygen is always −2 except in the compound OF2 and as part of peroxide when it is -1.
	+ Hydrogen is always +1 except in metal hydrides.
	+ Alkali (group I) metals are always +1.
	+ Alkaline Earth (group II) metals are always +2.
	+ Al is always +3, Zn is always +2, and Ag is always +1.

**Steps to solving oxidation states for the elements in a compound:**

1. Identify the type of molecule: elemental, covalent, ionic.
2. Figure out which elements’ oxidation states are known.
3. Work your way backwards to calculate the oxidation states of the other elements which can take on multiple charges.

For example, in the compound Na2HPO4:

* + - Na2HPO4 is
		- Your knowns are
* Sodium is group 1: so the oxidation number of Na is \_\_\_\_\_\_
* Oxygen is (almost) always \_\_\_\_\_.
* Hydrogen is not a hydride in this compound: so the oxidation number of H is \_\_\_
	+ - The HPO42− ion has a charge of \_\_\_\_\_. This means the oxidation numbers of H, P, and O must add up to \_\_\_\_.
			* + O = −2. There are 4 O atoms, so the negative oxidation numbers add up to \_\_\_\_.
				+ H is \_\_\_\_.
				+ If the O atoms add up to −8 and H is +1, then P must be \_\_\_\_\_.

Practice: Determine the oxidation states for the elements in the following compounds:

1. H2SO4
2. H2SO3
3. KH(NO3)2
4. AlPO3

Try It! Put some time in with the oxidation state practice worksheet.

**Balancing REDOX Reactions**

To fully balance a redox reaction, you must balance:

Often, redox reactions are shown and balanced as \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_. In this case, balancing them is often a simple matter of making sure that the same number of electrons are produced by the oxidation half-reaction and consumed by the reduction half-reaction.

For example, consider the unbalanced net ionic equation:

Al(s) + Zn2+(aq) → Al3+(aq) + Zn(s)

In this reaction, Al is \_\_\_\_\_\_\_\_\_\_\_\_\_\_ from \_\_\_\_\_ to \_\_\_\_\_\_, and Zn is \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ from \_\_\_\_\_ to \_\_\_\_\_. The *atoms* appear balanced, but Zn2+ needs only 2 electrons to form Zn0, but Al0 produces 3 electrons when oxidized to Al3+.

Write out the two half reactions for the RedOx reaction:

To balance the electrons, we need to multiply the first half-reaction by 2, and the second one by 3, giving:

If we combine these and cancel the electrons (because we have the same number on both sides), we get the balanced net ionic equation:

Practice

Write out the half reactions for the following redox reactions, balance them, and write the final net ionic equation for the reactions.

1. Cu2+(aq) + Al(s) Al3+(aq) + Cu(s)
2. Br-(aq) + Cl2(g) Cl-(aq) + Br2(l)

**Balancing RedOx Reactions in Acidic or Basic Solutions.**

In some redox reactions that take place in acid or base, H+ or OH− ions from the solution participate in the redox reaction. This makes balancing them a little more complicated.

In these reactions, in addition to balancing atoms and electrons, we also need to balance the total positive and negative \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ on both sides, using H+ or OH− ions from the solution and forming H2O.

In order to keep track of everything, I use the mnemonic AEIOU.

**A** = **Atoms**. Balance the atoms (except for H and O) in each half-reaction.

**E** = **Electrons**. Add the electrons gained or lost to each half-reaction.

**I** = **Ions**. Add H+ (if the reaction is in acid) or OH− (if the reaction is in base) to balance the charges on each side of each half-reaction.

**O** = **Oxygen** & **Hydrogen**. Balance oxygen and hydrogen in each half-reaction by adding H2O molecules as needed.

**U** = **Unite**. Unite the two half-reactions by multiplying all coefficients in each one by the factor needed to make the electrons in one cancel the electrons in the other.

Then cancel anything that appears on both sides to get the balanced net ionic equation.

For example:

Balance the following reaction in basic solution:

Ca (s) + Cr2O72− (aq) → CaO (s) + Cr3+ (aq)

The half-reactions are:

**Atoms** in each half-reaction:

The first half-reaction is already balanced for calcium atoms. (We’ll worry about oxygen later.)

Ca0 → CaO

We need 2 Cr3+ ions to balance the 2 Cr atoms in Cr2O72−:

Cr2O72− → 2 Cr3+

**Electrons** in each half-reaction:

Ca is being oxidized from Ca0 to Ca+2. This produces \_\_\_\_ electrons:

Ca0 → Ca+2O−2 + \_\_\_\_\_

Cr is being reduced from Cr+6 in Cr2O72− to Cr3+. This consumes 3 electrons per Cr atom, or \_\_\_\_\_ electrons total:

Cr2+6O72− + \_\_\_\_\_ → 2 Cr3+

**Ions** of H+ or OH− in each half-reaction to balance charges.

The total charge (ions and electrons) in the first half-reaction is 0 on the left and −2 on the right.

Ca0 → CaO + 2 *e−*

We have OH− ions to work with, so we add \_\_\_\_\_ of them on the left to bring the total charge on both sides to −2:

Ca0 + \_\_\_\_\_\_\_ → CaO + 2 *e−*

In the second half-reaction, the total charge is −8 on the left and +6 on the right:

Cr2O72− + 6 *e−* → 2 Cr3+

We add \_\_\_\_\_\_\_\_\_ ions to the right to bring the total charge on both sides to −8:

Cr2O72− + 6 *e−* → 2 Cr3+ + \_\_\_\_\_\_\_\_\_

**Oxygen** & **Hydrogen** atoms in each half-reaction:

In the first half-reaction, there are 2 unbalanced H atoms and one unbalanced O atom on the left.

Ca0 + 2 OH− → CaO + 2 *e−*

To balance these, we need \_\_\_\_\_ H2O molecule on the right:

Ca0 + 2 OH− → CaO + 2 *e−* + \_\_\_\_\_\_\_\_\_

The second half-reaction has 14 unbalanced H atoms and 7 unbalanced O atoms on the right.

Cr2O72− + 6 *e−* → 2 Cr3+ + 14 OH−

To balance these, we need \_\_\_\_\_ H2O molecules on the left:

\_\_\_\_\_\_\_ + Cr2O72− + 6 *e−* → 2 Cr3+ + 14 OH−

**Unite** the half-reactions with equal numbers of electrons produced/ consumed:

The first half-reaction produces 2 electrons…

Ca0 + 2 OH− → CaO + 2 *e−* + H2O

and the second half-reaction consumes 6 electrons…

7 H2O + Cr2O72− + 6 *e−* → 2 Cr3+ + 14 OH−

This means we have to multiply the first one by 3 to balance the electrons:

3(Ca0 + 2 OH− → CaO + 2 *e−* + H2O)

\_\_\_\_Ca0 + \_\_\_\_OH− → \_\_\_\_CaO + \_\_\_\_*e−* + \_\_\_\_H2O

Adding the two half-reactions together, we get:

Finally, we cancel anything that appears on both sides. We can cancel:

The final, balanced redox reaction in net ionic form is: