Oxidation/Reduction

When you studied the topic of atomic structure you learned that:
Protons have a + 1 charge
Electrons have a - 1 charge.
All atoms are neutral (no charge).

What is the charge of a proton? _____

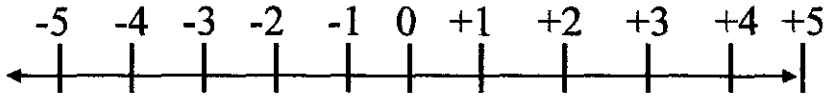
What is the charge of an electron? _____

What is the charge of any atom? _____

All atoms are neutral because an atom has = amounts of + **protons** and - **electrons.**
The result is a **neutral (no charge atom).**

Why are all atoms neutral? _____
If an atom gains or loses electrons, it is no longer neutral and can no longer be called an atom. Instead, it is called an **ion (charged particle).**
An **oxidation state (number)** is the charge of an ion or the apparent charge of an atom during bonding.
Oxidation #s are 0 in the elemental form. Exs: $Mg^0_{(s)}$, $H^0_{2(g)}$, & $Br^0_{(2)}$

Understanding a number scale:
Caution: A gain or loss of one or more electrons may result in the one or more of the following:
a) A negative oxidation # can become more or less negative.
b) A positive oxidation # can become more or less positive.
c) A negative oxidation # can become positive.
d) A positive oxidation # can become negative.
e) A neutral "0" oxidation # can become positive or negative.



A gain of electrons will move the value to the left. ←
A loss of electrons will move the value to the right. →

Ex: If a particle with a charge of -1 gains 2 electrons, what will be the resulting charge?
Result: - 3.

Ex: If a particle with a charge of -1 loses 2 electrons, what will be the resulting charge?
Result: + 1.

a) If an ion with a + 2 charge gains 3 electrons, what will be the result?

b) If a neutral atom loses 1 electron, what will be the result?

Oxidation refers to the process of a particle (atom or ion) **losing** one or more **electrons. Oxidation results in an increase in oxidation #.** Electrons have a - 1 charge, every time a particle loses an electron its charge increases by 1. If a particle loses 2 electrons its charge increases by 2.

What does oxidation refer to? _____

What happens to an oxidation # during oxidation?_____

Ex: If an aluminum atom loses 3 electrons, what will be the result?
Result: An aluminum ion with a + 3 charge.

Ex: If a Fe^{+2} ion loses an electron, what will be the result?
Result: a Fe^{+3} ion.

a) If a zirconium atom loses 4 electrons, what will be the result?

b) If a Pb^{+2} ion loses 2 electrons, what will be the result?

Reduction refers to the process of a particle (atom or ion) **gaining** one or more **electrons. Reduction results in a decrease (reduction) of oxidation #.** Electrons have a - 1 charge, every time a particle gains 1 electron its charge decreases by 1. If a particle gains 2 electrons its charge decreases by 2.

What does reduction refer to? _____

What happens to an oxidation # during reduction? _____

Ex: If a sulfur atom gains 2 electrons it will become a sulfide ion with a - 2 charge. It can be shown as S^{-2}.

Ex: If a Cu^{+1} ion gains an electron, what will be the result?
Result: a Cu^{0} atom.

a) If a N^{-3} ion gains an electron what will be the result?

b) If a Mg^{+2} ion gains 2 electrons, what will be the result?

Lose electrons oxidation. **Leo**
Gain electrons reduction. **Ger**

Leo the Lion is here to help you remember the difference between oxidation and reduction.

Rules for Determining Oxidation Numbers for Elements in Compounds:
You have already learned some things about oxidation numbers. **The sum of the oxidation numbers in a compound must equal zero.** The charge on an ion is determined by adding the charges of the + protons to the charges of the - electrons.

protons – electrons = charge

Important! If an atom of an element is not combined with another element it is said to be in the free or elemental state. **The oxidation # of any atom in the free state is 0.** This includes diatomics and allotropes of elements. Exs: Br_2, O_3 & S_8.
Examples: The oxidation # of H in H_2 is 0. The oxidation # of Cl in Cl_2 is 0.
You may see the elemental form of an element expressed as: Mg^0
Some elements have only 1 possible oxidation #. Others have many choices.
Oxidation numbers can be found in the periodic table.

Oxidation # exceptions:

a) Hydrogen is always + 1 unless it is combined in a compound with a Group 1 or Group 2 metal. Ex: KH, NaH, MgH_2 & CaH_2. If so, hydrogen will have a - 1 oxidation #.

b) Oxygen is almost always – 2.
However, if oxygen is combined with F as OF_2 , oxygen takes on a + 2 oxidation #.

Also, When oxygen is in the polyatomic ion "peroxide" $(O_2)^{-2}$, oxygen will have a - 1 oxidation #.

Determining Oxidation #s

The sum of the oxidation # s in a compound is 0.

What is the sum of the oxidation # s in a compound? _____

Ex: NaCl Na is +1 it has no other choice. Cl has to be - 1 to make the compound 0.

Ex: NO_2 O can only be -2 therefore N must be + 4.

Try one! What is the oxidation # of each element in Al_2S_3?

What do you do when there are more than 2 elements present in a compound?
Insert the oxidation #s of the elements that only have only one possible oxidation #. Then calculate the value of the remaining one.

Ex: What is the oxidation # of each element in $KClO_3$?
To solve: Notice there is 1 K , 1Cl, and 3 Os.
K can only be + 1.
O can only be -2 and there are 3 of them. 3 (- 2) = -6.
Cl has more than 1 choice from the periodic table. Its value must bring the total to zero.
+ 1 –6 + ? = 0 Cl must have a value of + 5.

a) What is the oxidation # of N in KNO_3?

b) What is the oxidation # of P in $AlPO_4$?

c) What is the oxidation # of Cl in $NaClO_4$?

The sum of the oxidation # s in a polyatomic ion is equal to the charge of the polyatomic ion.

To determine the oxidation number of each element in a polyatomic ion:

Insert the oxidation #s of the elements that only have only one possible oxidation #.
Then calculate the value of the remaining one.

Ex: What is the oxidation # of each element in ClO_3^{-1}?
To solve: Notice there is 1Cl, and 3 Os. The charge of the polyatomic ion is -1.
O can only be -2 and there are 3 of them. 3 (- 2) = -6.
Cl has more than 1 choice from the periodic table. Its value must bring the total to -1.
$-6 + \underline{?} = -1$ Cl must have a value of **+ 5**.

a) What is the oxidation # of N in NO_3^{-1}?

b) What is the oxidation # of Cr in $Cr_2O_7^{-2}$?

c) What is the oxidation # of N in NH_4^{+1}?

Redox is a term that describes the process of **reduction** and **oxidation** occurring at the same time (simultaneously). Some particles (atoms or ions) lose one or more electrons. Some other particles (atoms or ions) gain those lost electrons.
Why does redox occur? Atoms compete for electrons. Remember the terms **electronegativity** and **ionization energy?** I didn't think so, it was along time ago.
Electronegativity is an atom's attraction for electrons.

Ionization energy is the amount of energy required to remove an atom's outermost electron.

Describe: electronegativity _____

_____

Describe: ionization energy _____

_____

These 2 factors combine to determine which atoms will gain or lose electrons. Electronegativity and ionization energy values can be found on pages 10 & 11 of the reference table.

Table J is used to determine which atoms will gain or lose electrons in a redox reaction and can be used to predict if single replacement reactions will occcur.

Elements towards the top of Table J have low ionization energies and low electronegativities. Elements towards the top of the table readily lose electrons. Elements towards the bottom of the table have higher electronegativities and ionization energies. They tend to gain electrons from the elements above them in the table. Remember! In the topic of acids and bases Table J was also used to predict which metals would react with acids in a single replacement reaction to produce hydrogen.

Elements towards the top of Table J readily _____ electrons.

Elements towards the bottom of table J readily _____ electrons.

Redox Reactions: For a redox reaction to occur oxidation and reduction must occur at the same time. Something will lose electrons and something will gain those electrons. Redox reactions obey the Law of Conservation of Mass, the Law of Conservation of charge, and the law of Conservation of Energy.

Identifying Redox Reactions:
If a reaction is redox, some item in the reaction must have a change of oxidation #.
Ex: $2H_{2(g)} + O_{2(g)} \rightarrow 2H_2O_{(l)}$
Hydrogen as a reactant is not combined with any other element. It is a free element. Its oxidation # is 0. As a product hydrogen is part of a compound. In the compound its oxidation # is + 1.
Oxygen as a reactant is not combined with any other element. It is a free element. Its oxidation # is 0. As a product oxygen is part of a compound. In the compound its oxidation # is - 2.
The above reaction is a redox reaction both hydrogen and oxygen experienced a change in oxidation #.

Ex: Is this reaction redox? **$4\ Fe_{(s)} + 3\ O_{2(g)} \rightarrow 2\ Fe_2O_{3(aq)}$** Why?
Iron as a reactant is not combined with any other element. It is a free element. Its oxidation # is 0. As a product Iron is part of a compound. In the compound its oxidation # is + 3.
Oxygen as a reactant is not combined with any other element. It is a free element. Its oxidation # is 0. As a product oxygen is part of a compound. In the compound its oxidation # is - 2.
The above reaction is a redox reaction because both iron and oxygen experienced a change in oxidation #.

Ex: Is the following reaction redox? **HCl + NaOH → NaCl + HOH** Why?
Solve: Determine the oxidation #s of all the elements.

$H^{+1}\ Cl^{-1} + Na^{+1}\ O^{-2}\ H^{+1} \rightarrow Na^{+1}\ Cl^{-1} + H^{+1}\ O^{-2}\ H^{+1}$
Notice: nothing changed oxidation #. Therefore; no redox reaction

Short Cut! An easy way to determine if a reaction is redox is to see if any element is a free element on one side of the → and combined in a compound on the other side of the →. If it is, it is redox.
Why? Because if any element is a free element its oxidation # is 0. If it joins a compound it will have to have a # other than zero. Therefore, there must have been a change in oxidation #.
Ex: $Mg_{(s)} + 2\ HCl_{(aq)} \rightarrow MgCl_{2(aq)} + H_{2(g)}$
Notice: Mg as a reactant is a free element. As a product Mg is combined in a compound.
Therefore; Mg must have had a change in oxidation #.
Therefore; the reaction is redox.

a) Is the following reaction redox? **Why?** $Zn_{(s)} + H_2SO_{4(aq)} \rightarrow ZnSO_{4(aq)} + H_{2(g)}$

b) Is the following reaction redox? **Why?**

$2\ Al_{(s)} + 3\ H_2SO_{4(aq)} \rightarrow Al_2(SO_4)_{3(aq)} + 3H_{2(g)}$

c) Is the following reaction redox? **Why?**

$Ag(NO_3) + NaCl \rightarrow AgCl + Na(NO_3)$

d) Is the following reaction redox? **Why?**

$Mg + CuSO_4 \rightarrow MgSO_4 + Cu$

Use Table J to Predict if a single replacement redox reaction will occur.
You have already had experience using **Table J** in the topic of Acid/base.
The topic of **redox** uses Table J to predict whether a single replacement reaction will occur.
Rule: The closer an element is towards the top of the table the more active it is. This means that **elements towards the top will lose electrons (oxidation).** Active elements will form ions or compounds. Active elements are usually not found in the free (elemental) state in nature.
An active element will replace a less active element in a **single replacement reaction.** Remember those?

Ex: $Zn_{(s)} + HCl_{(aq)} \rightarrow ZnCl_{2(aq)} + H_{2(g)}$
Notice: Zn is higher on the table, therefore, it replaced H in a single replacement reaction. Zn was oxidized and H was reduced.

Predict the results of the following reactions. **Write proper formulas and balance.**

1. $Mg_{(s)} + AlCl_{3(aq)} \rightarrow$

2. $Ca_{(s)} + BaCl_{2(aq)} \rightarrow$

3. $Ag_{(s)} + CuCl_{2(aq)} \rightarrow$

4. $Na_{(s)} + HOH_{(l)} \rightarrow$

5. $Fe_{(s)} + MgBr_{2(aq)} \rightarrow$

Half Reactions

All redox reactions experience both **red**uction and **ox**idation. All redox reactions consist of 2 parts. One part is a gain of electrons (reduction) **"GER"** .
The other part is a loss of electrons (oxidation) **"LEO"**.
A redox reaction can be separated into an oxidation half and a reduction half.

A separate equation showing electrons lost can be written.
A separate equation showing electrons gained can be written.

To properly write a half reaction you must remember how to assign oxidation #s.

Half reactions must obey the Laws of Conservation of Mass and Charge!

Rule:
An **oxidation half reaction shows electrons as products** (to the right of the arrow).

A **reduction half reaction shows electrons as reactants** (to the left of the arrow). **Re**duction **re**actants

Look at the following redox reaction.

$Mg + Cl_2 \rightarrow MgCl_2$

oxidation half reaction: $Mg^0 \rightarrow Mg^{+2} + 2\ e^-$

reduction half reaction: $Cl_2 + 2e^- \rightarrow 2Cl^-$

Ex: $Ca + Cu^{+2} \rightarrow Ca^{+2} + Cu$

oxidation half reaction: $Ca \rightarrow Ca^{+2} + 2\ e^-$

reduction half reaction: $Cu^{+2} + 2e^- \rightarrow Cu$

The sum of the charges on both sides of the arrow must be equal.

The numbers of each element on each side of the arrow must be equal.

Write half reactions for the following redox reactions and indicate which one is **oxidation** and **reduction**.

1. $Mg + Zn^{+2} \rightarrow Mg^{+2} + Zn$

2. $Fe + Zn^{+2} \rightarrow Fe^{+2} + Zn$

3. $2Al + 3Fe^{+2} \rightarrow 2Al^{+3} + 3Fe$

4. $Zn_{(s)} + Cl_{2(aq)} \rightarrow ZnCl_{2(aq)}$

5. $2H_2 + O_2 \rightarrow 2H_2O$

6. $Mg + ZnS \rightarrow MgS + Zn$

Balancing Redox Reactions:

1. Redox reactions obey the Law of Conservation of Mass.
   The number of each type of atom must be the same on both sides of the arrow.
2. Redox reactions obey the Law of Conservation of Charge.
   The number of electrons lost must equal the number of electrons gained.

Balancing redox reactions may look complicated but if you follow the method they are not so bad.

Method:

1) Use the oxidation # of the first reactant as the coefficient of the second reactant. Use the oxidation# of the second reactant as the coefficient of the first reactant. If you don't see an element's oxidation #, use its oxidation # from the periodic table.
2) Use the oxidation # of the first product from as the coefficient of the second product. Use the oxidation # of the second product as the coefficient of the first product. If you don't see an oxidation #, use the oxidation # from the periodic table.

Lets start with an example.

Balance this redox reaction: $\_\_Hg^{+2} + \_\_Ag \rightarrow \_\_Hg + \_\_Ag^{+}$

$\underline{1}\,Hg^{+2} + \underline{2}\,Ag \rightarrow \underline{1}\,Hg + \underline{2}\,Ag^{+}$

a) $\_\_Fe + \_\_Ag^{+} \rightarrow \_\_Ag + \_\_Fe^{+3}$

b) $\_\_Cr^{+3} + \_\_Ca \rightarrow \_\_Cr + \_\_Ca^{+2}$

c) $\_\_\_Mg + \_\_\_Fe^{+3} \rightarrow \_\_\_Mg^{+2} + \_\_\_Fe$