General Chemistry

202-SN1-RE

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Unit 3 – Types of Chemical Reactions

Chemical Reactions

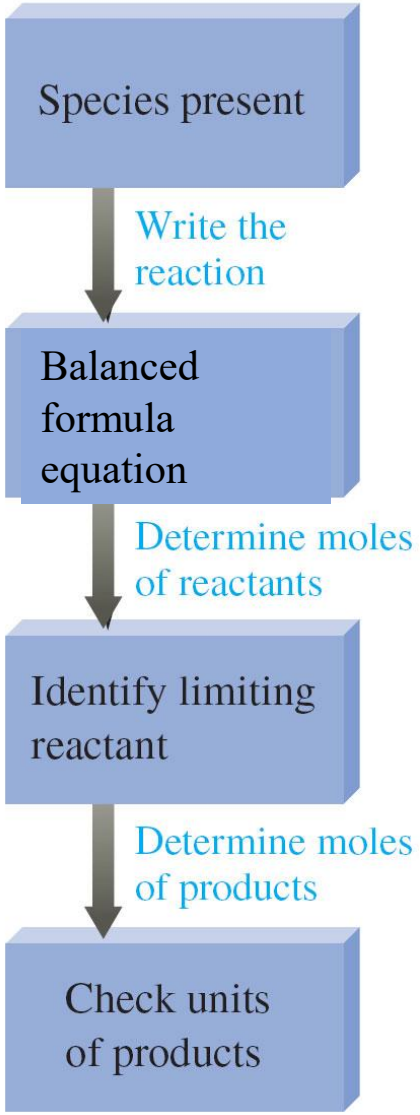
- In this course, we will be discussing the following types of reactions:
  - **Precipitation**
    - Formation of an insoluble solid (precipitate)
  - **Acid-Base**
    - Formation of a salt and water
  - **Oxidation-Reduction**
    - Transfer of electrons from a reductant (electron donor) to an oxidant (electron acceptor)

Stoichiometry for Reactions in Solutions

Stoichiometry for Reactions in Solutions

Steps to solving stoichiometry problems for reactions in solution:

1. Identify the species present in the combined solution, and determine what reaction if any occurs.
2. Write the balanced formula equation for the reaction.
3. Calculate the moles of reactants.
4. Determine which reactant is limiting.
5. Calculate the moles of product(s), as required.
6. Convert to grams or other units, as required.



Acid-Base Reactions

Acid-Base Reactions

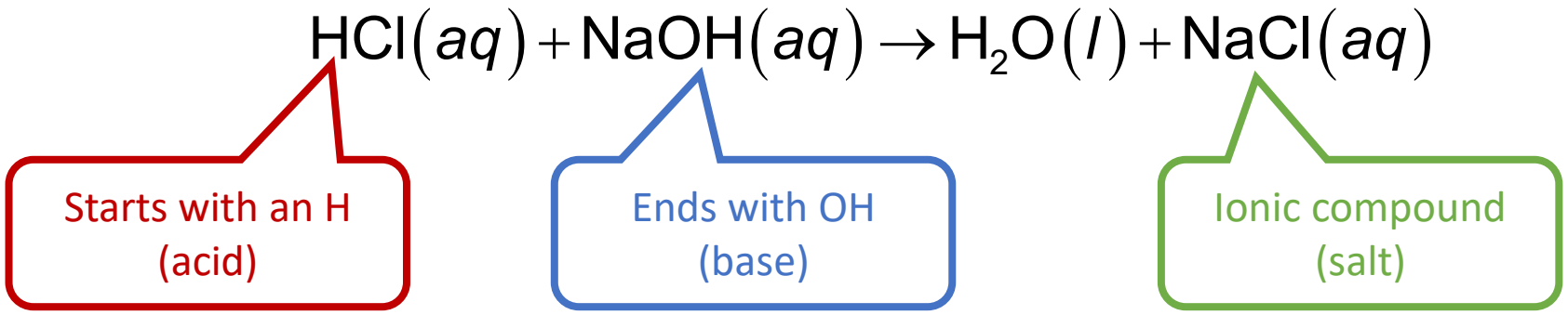
- There is a strong tendency for a reaction to occur if the substances mixed can form water.

- Acid-base reactions can be summarized as a proton (H^+) transfer from an acid to a base.

$$HCl_{(aq)} + NaOH_{(s)} \rightarrow NaCl_{(aq)} + H_2O_{(l)}$$

Acid-Base Neutralization Reactions

Acid–Base Neutralization Reactions: Processes in which an acid reacts with a base to yield water plus an ionic compound called a salt.



Example: Write the molecular equation of the reaction between hydroiodic acid and barium hydroxide.

Identify the compounds:

$$HI(aq) + Ba(OH)_2(aq) \longrightarrow H_2O(l) + BaI_2(aq)$$

acid base water salt

Balanced molecular equation:

$$2\,HI(aq) + Ba(OH)_2(aq) \longrightarrow 2\,H_2O(l) + BaI_2(aq)$$

Models of Acids and Bases

- Arrhenius: Acids produce H^+ ions in solution, bases produce OH^- ions.

$$HCl \xrightarrow{H_2O} H^+(aq) + Cl^-(aq)$$

$$NaOH(s) \xrightarrow{H_2O} Na^+(aq) + OH^-(aq)$$

- Brønsted–Lowry: Acids are proton (H^+) donors, bases are proton acceptors.

Eg. $HCl + H_2O \rightarrow Cl^- + H_3O^+$

acid base

Characteristics of an Acid-Base Reaction

- When to classify a reaction as an acid-base reaction:
  - Strong acids and strong bases are strong electrolytes. They will always give the following net ionic equation:

$$HCl(aq) + KOH(aq) \rightarrow H_2O(l) + KCl(aq)$$

$$H^+(aq) + \cancel{Cl^-(aq)} + \cancel{K^+(aq)} + OH^-(aq) \rightarrow H_2O(l) + \cancel{K^+(aq)} + \cancel{Cl^-(aq)}$$

$$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$$

  - Can also have formation of a gas in an acid-base reaction.

$$2HCl(aq) + Na_2CO_3(aq) \rightarrow CO_2(g) + H_2O(l) + 2\ NaCl(aq)$$

Net ionic equation: $2H^+(aq) + CO_3^{2-}(aq) \rightarrow CO_2(g) + H_2O(l)$

Combining Stoichiometry and Aqueous Solutions

What volume (in mL) of 0.250 M sulfuric acid is needed to react with 50.0 mL of 0.100 M sodium hydroxide?

$$H_2SO_4(aq) + 2NaOH(aq) \rightarrow Na_2SO_4(aq) + 2H_2O(l)$$

Moles of NaOH available:

$$50.0 \text{ mL NaOH} \times \frac{0.100 \text{ mol}}{1 \text{L}} \times \frac{1 \text{ L}}{1000 \text{ mL}} = 0.00500 \text{ mol NaOH}$$

Volume of H_2SO_4 needed:

$$0.00500 \text{ moL NaOH} \times \frac{1 \text{ mol } H_2SO_4}{2 \text{ mol NaOH}} \times \frac{1 \text{ L solution}}{0.250 \text{ mol } H_2SO_4}$$

10.0 mL solution (0.250 M H_2SO_4)

Oxidation-Reduction Reactions

Oxidation-Reduction (Redox) Reactions

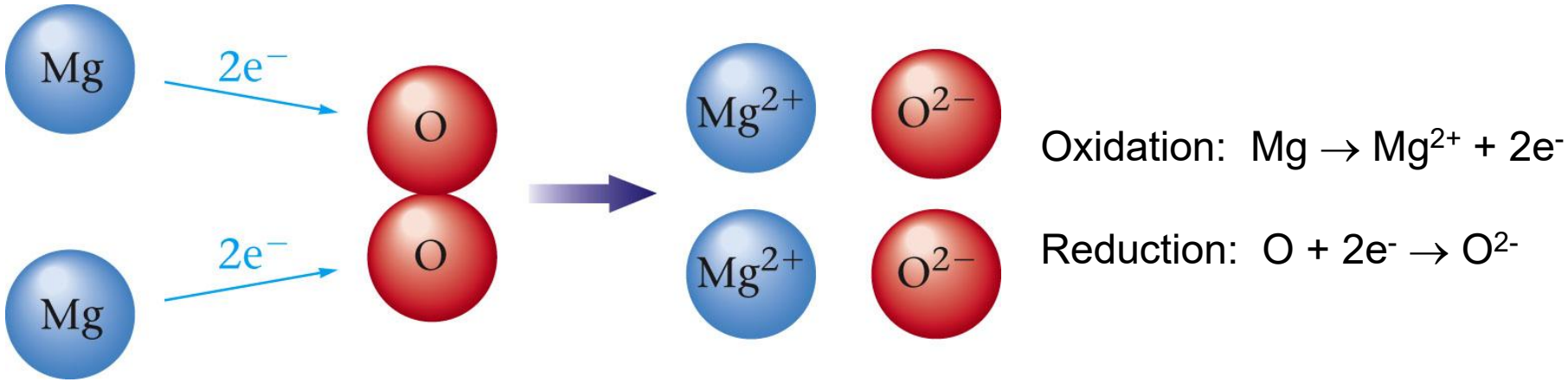
Traditionally, oxidation meant combining with oxygen to produce an oxide and reduction meant removing oxygen to get an element.

Oxidation of Fe (rusting) $4Fe(s) + 3O_2(g) \rightarrow 2Fe_2O_3(s)$

Reduction of Fe_2O_3 (making iron) $2Fe_2O_3(s) + 3C(s) \rightarrow 4Fe(s) + 3CO_2(g)$

Oxidation-Reduction Reactions (Redox)

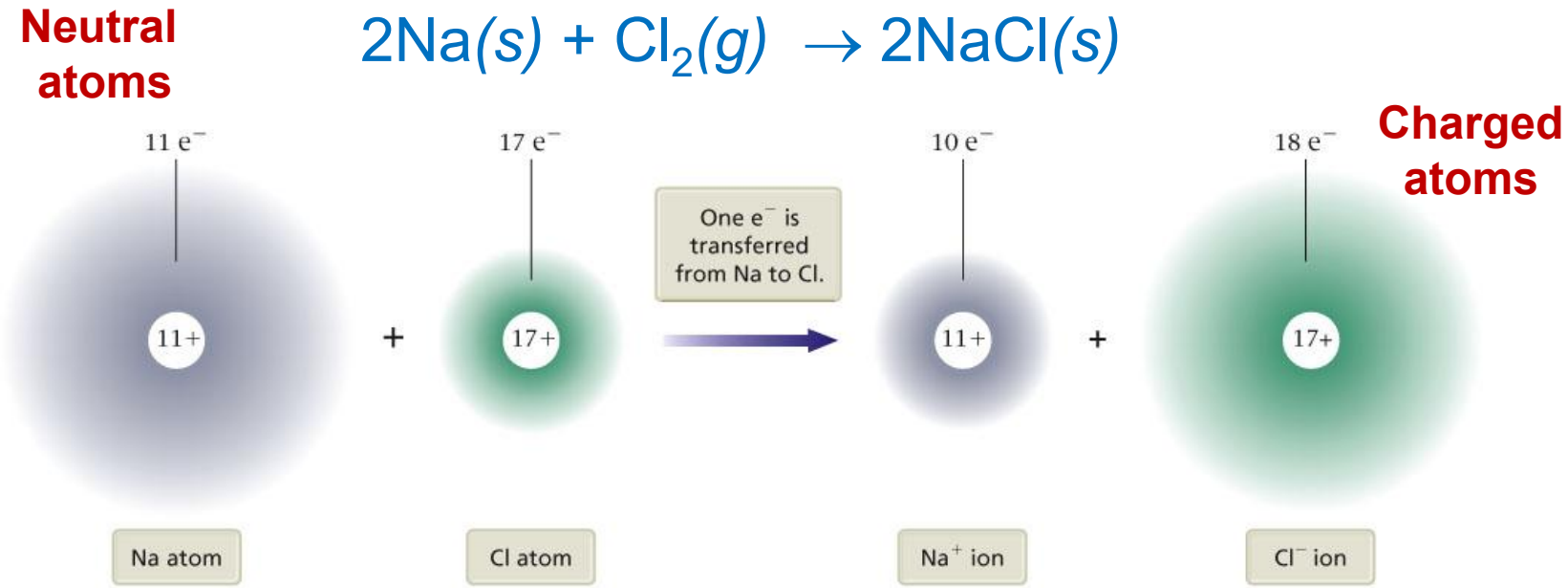
- A reaction that involves a transfer of electrons from one reactant to another
- Two simultaneous reactions occurring:
  - **Oxidation**: Loss/donation of electrons
  - **Reduction**: Gain/acceptance of electrons
- Ex: $2Mg(s) + O_2(g) \rightarrow 2MgO(s)$



Oxidation-Reduction Reactions

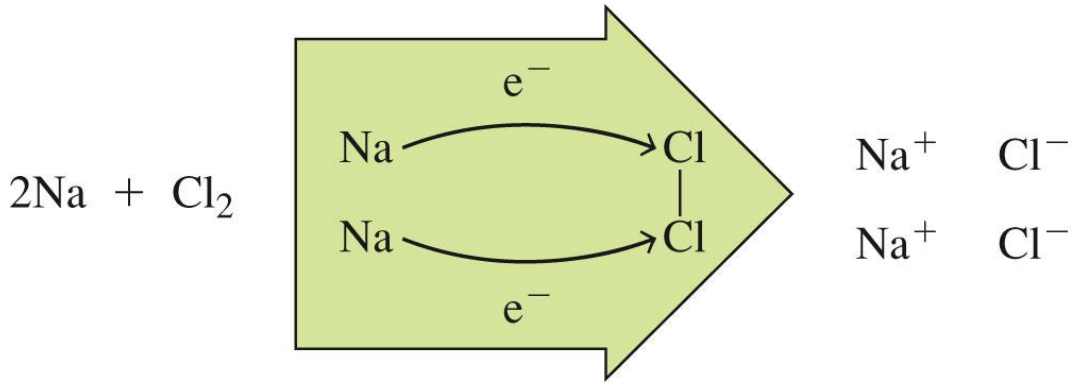
- Oxidation-reduction reaction between a metal and nonmetal forms an ionic compound.

- Electrons are always transferred from the metal (which forms a cation) to a nonmetal (which forms an anion).



Oxidation-Reduction Reactions

- Oxidation and reduction reactions must take place simultaneously in a system. Also called a **redox** reaction.



- Handy mnemonics:
  - **O**xidation **i**s the **l**oss of electrons
  - **R**eduction **i**s the **g**ain of electrons

  } OIL RIG

  - **L**oss of **e**lectrons, **o**xidation
  - **G**ain of **e**lectrons, **r**eduction

  } GER says LEO

Oxidation-Reduction (Redox) Reactions

Question 4.8 Is the following reaction (in the forward direction) an oxidation or a reduction? Support your choice.

$$Cu^{2+} + e^- \longrightarrow Cu^+$$

a) **Oxidation** because an electron is **gained** by the Cu^{2+} ion.

b) **Oxidation** because an electron is **lost** by the Cu^{2+} ion.

c) **Reduction** because an electron is **gained** by the Cu^{2+} ion.

d) **Reduction** because an **electron** is lost by the Cu^{2+} ion.

Oxidation-Reduction (Redox) Reactions

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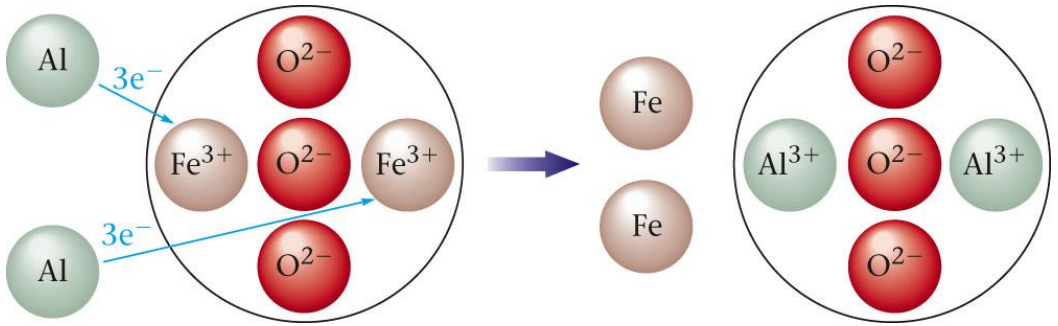
Oxidation-Reduction Reactions

Which of the following best describes what is happening in the following representation of an oxidation–reduction reaction:



a) Metal Al gains 3 e^- and O^{2-} in Fe_2O_3 loses these $3e^-$.

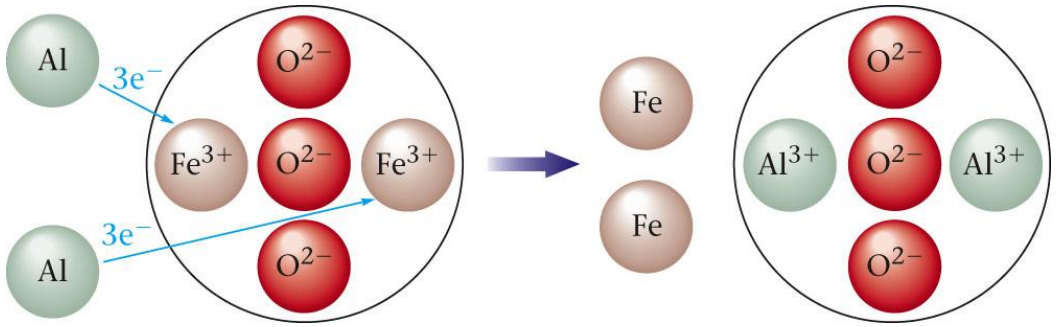
b) Metal Al gains 3 e^- and Fe^{3+} in Fe_2O_3 loses these $3e^-$.

c) Metal Al loses 3 e^- and O^{2-} in Fe_2O_3 gains these $3e^-$.

d) Metal Al loses 3 e^- and Fe^{3+} in Fe_2O_3 gains these $3e^-$.

Oxidation-Reduction Reactions

Which of the following best describes what is happening in the following representation of an oxidation–reduction reaction:



a) Metal Al gains 3 e^- and O^{2-} in Fe_2O_3 loses these $3e^-$.

b) Metal Al gains 3 e^- and Fe^{3+} in Fe_2O_3 loses these $3e^-$.

c) Metal Al loses 3 e^- and O^{2-} in Fe_2O_3 gains these $3e^-$.

d) Metal Al loses 3 e^- and Fe^{3+} in Fe_2O_3 gains these $3e^-$.

Oxidation States

- Defining the oxidation state (or oxidation number) of an atom provides a way to keep track of electrons in redox reactions (especially in covalent substances).

- It's a way to see which atoms are losing electrons and which atoms are gaining electrons.

- Compare ionic vs covalent substances:

$$2Na(s) + Cl_2(g) \rightarrow 2NaCl(s)$$

Electrons are transferred to form ions.

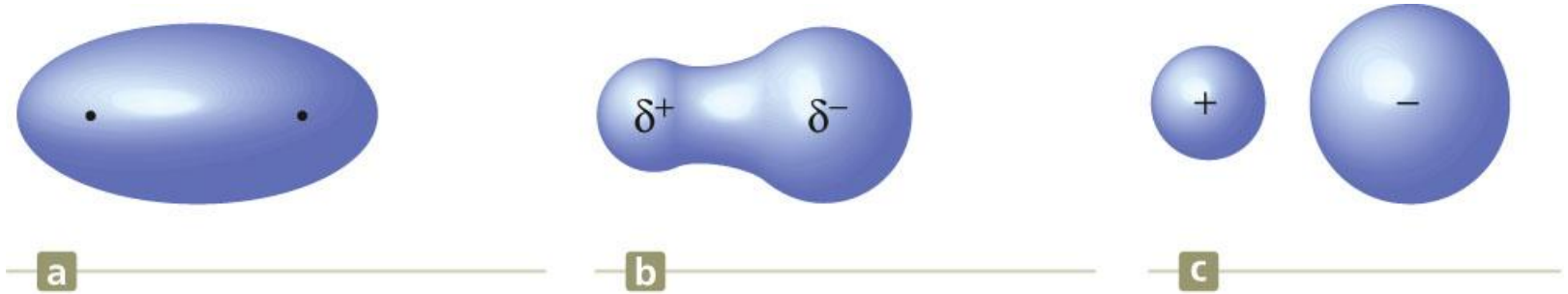
$$C_7H_{12} + 11\ O_2 \rightarrow 7\ CO_2 + 8\ H_2O$$

Ions are not formed, but electrons have been transferred.

- Oxidation numbers are similar to assigning charges to the atoms in a compound.

Electronegativity

- What is it?
  - A periodic trend
  - It is a description of how much an atom wants electrons.
  - The higher the electronegativity, the higher the want.



A covalent bond formed between identical atoms.

Equal electronegativity

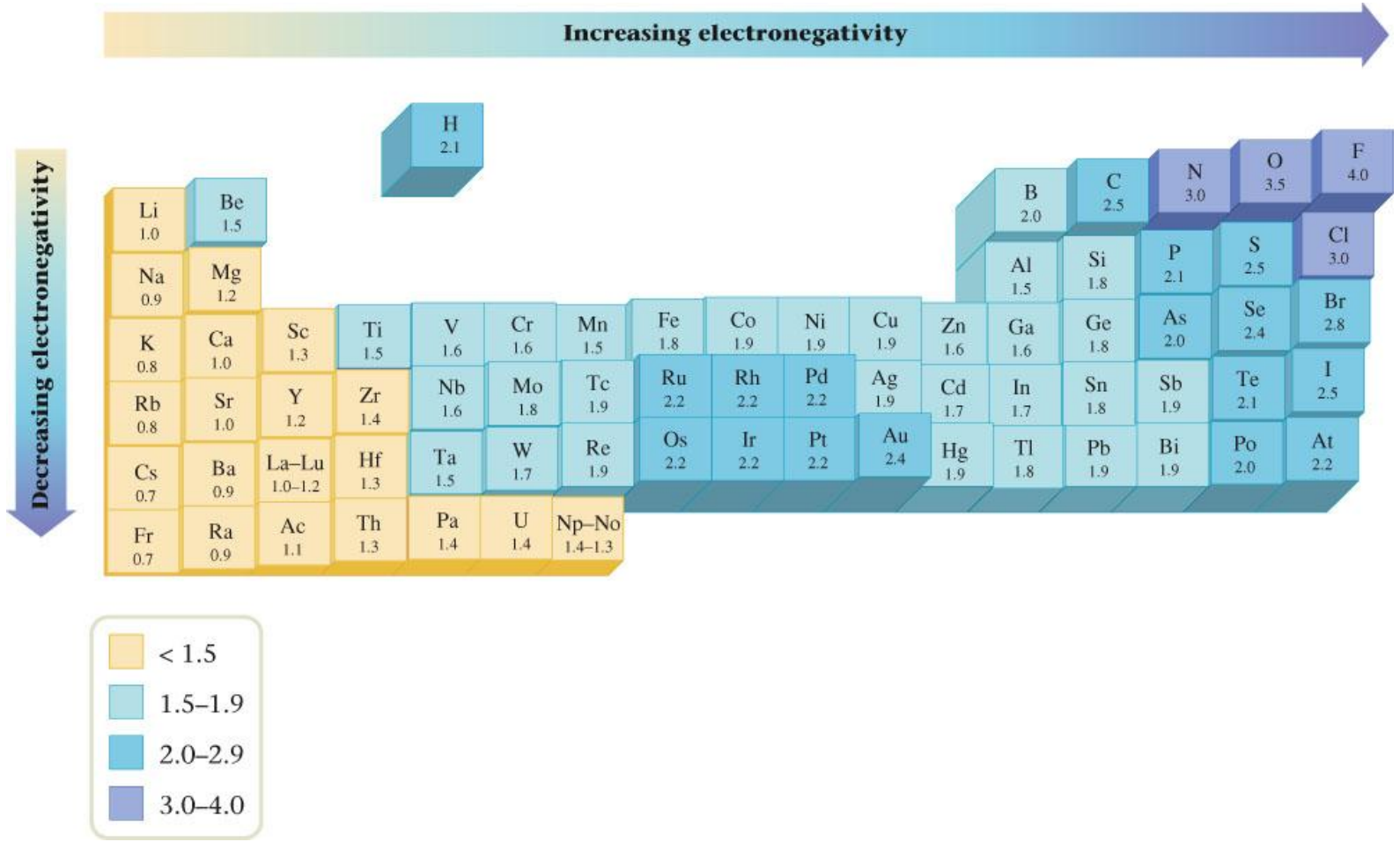
A polar covalent bond, with both ionic and covalent components.

Slightly unequal electronegativity

An ionic bond, with no electron sharing.

Drastically different electronegativity

Electronegativity



Rules	Oxidation State	Examples
atom in an element	0	Na(s), O_2(g), Hg(l), S_8(s)
monatomic ion	charge of the ion	Al^{3+} would be +3 Br^- would be -1
oxygen	-2 (unless it's a peroxide, in which case -1)	O in H_2O would be -2 O in H_2O_2 (peroxide) would be -1
Hydrogen	+1 in covalent compounds -1 in ionic compounds	H in CH_4 would be +1 H in NaH would be -1
Fluorine	-1	F in KF would be -1 F in PF_3 would be -1
All Halogens (Group 7A elements)	-1 +n if bonded atom is more electronegative	F, Cl, Br, I, At Br in HBrO would be +1 Cl in $HClO_4$ would be +7
All Group 1A elements	+1	Li, Na, K, Rb, Cs, Fr
All Group 2A elements	+2	Be, Mg, Ca, Sr, Ba, Ra
Sum of oxidation states in compounds	0	H_2O, C_2H_4, CO_2 all would have an overall sum of 0
Sum of oxidation states in ions	overall charge of the ion	OH^- would be -1 NH_4^+ would be +1 CO_3^{2-} would be -2

Oxidation-Reduction (Redox) Reactions

1) **Elements:** oxidation number is always ____ for atoms in their elemental state.

2) **Neutral molecule:** The sum of oxidation numbers is ___ .

- Calculate the oxidation number for carbon in carbon dioxide (neutral molecule):

CO_2

3) **Polyatomic ions:** The sum of oxidation numbers equals the _____.

- Calculate the oxidation number for chromium (Cr) in this polyatomic ion.

$Cr_2O_7^{2-}$

Oxidation-Reduction (Redox) Reactions

1) **Elements:** oxidation number is always **0** for atoms in their elemental state.

2) **Neutral molecule:** The sum of oxidation numbers is **0** .

- Calculate the oxidation number for carbon in carbon dioxide (neutral molecule):

CO_2

$1(x) + 2(-2) = 0$

$x = +4$

3) **Polyatomic ions:** The sum of oxidation numbers equals the **net charge**.

- Calculate the oxidation number for chromium (Cr) in this polyatomic ion.

$Cr_2O_7^{2-}$

$2(x) + 7(-2) = -2$ (net charge)

$2x - 14 = -2$

$2x = 12$

$x = +6$

Oxidation-Reduction (Redox) Reactions

4) Monatomic ion: oxidation number is the same as its _____.

- Halogens usually have an oxidation number of –1. However, when it is bonded to oxygen, it needs to be calculated. Calculate the oxidation number for chlorine in these two compounds:

Hydrogen chloride, HCl

Perchloric acid, $HClO_4$

Oxidation-Reduction (Redox) Reactions

4) Monatomic ion: oxidation number is the same as its **charge**.

- Halogens usually have an oxidation number of –1. However, when it is bonded to oxygen, it needs to be calculated. Calculate the oxidation number for chlorine in these two compounds:

Hydrogen chloride, HCl

1(+1) + 1(Cl) = 0

Cl = -1

Perchloric acid, $HClO_4$

1(+1) + Cl + 4(-2) = 0

Cl = +7

Oxidation-Reduction (Redox) Reactions

5) Hydrogen: oxidation number can be either _____

- Calculate the oxidation number for hydrogen in these two species.

$(OH)^-$ CaH_2

6) Oxygen: usually has an oxidation number of ___ *(except for peroxides)*.

- Calculate the oxidation number for oxygen in these two compounds:

H_2O or H – O – H Hydrogen peroxide, H_2O_2 or H – O – O – H

Oxidation-Reduction (Redox) Reactions

5) Hydrogen: oxidation number can be either **+1 or -1.**

- Calculate the oxidation number for hydrogen in these two species.

$(OH)^-$

$x - 2 = -1$

$x = +1$

CaH_2

$1(+2) + 2(x) + = 0$

$2 + 2x = 0$

$x = -1$

6) Oxygen: usually has an oxidation number of **−2** *(except for peroxides)*.

- Calculate the oxidation number for oxygen in these two compounds:

H_2O or H – O – H

+1 -2 +1

Hydrogen peroxide, H_2O_2 or H – O – O – H

+1 -1 -1 +1

Oxidation-Reduction (Redox) Reactions

Practice: Identify the changes in oxidation states of the following reaction:

$$4Fe + 3O_2 \rightarrow 2\ Fe_2O_3$$

Change in oxidation number:

Fe:_____

O:_____

Oxidation-Reduction (Redox) Reactions

Practice: Identify the changes in oxidation states of the following reaction:

$$4Fe + 3O_2 \rightarrow 2\ Fe_2O_3$$

Change in oxidation number: 0 0 → $2(x) + 3(-2) = 0$

$x = +3$

Fe: 0 to +3

O: 0 to -2

Exercise

Find the oxidation states for each of the elements in each of the following compounds:

- $K_2Cr_2O_7$
- CO_3^{2-}
- MnO_2
- PCl_5
- SF_4

Exercise

Find the oxidation states for each of the elements in each of the following compounds:

- $K_2Cr_2O_7$
- CO_3^{2-}
- MnO_2
- PCl_5
- SF_4

K = +1; Cr = +6; O = –2

C = +4; O = –2

Mn = +4; O = –2

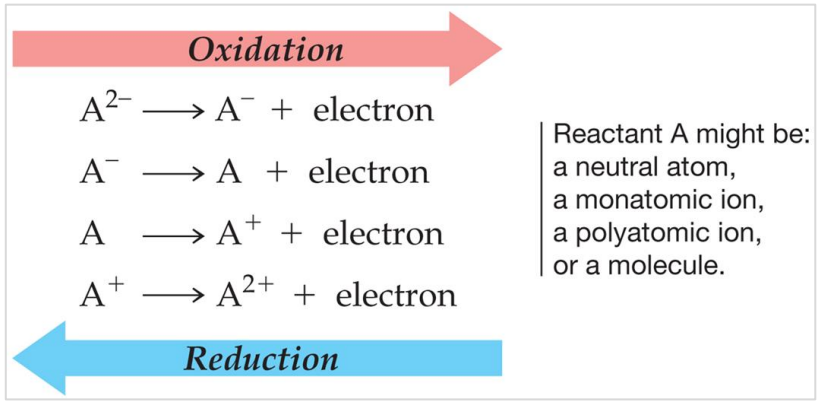
P = +5; Cl = –1

S = +4; F = –1

Oxidation-Reduction (Redox) Reactions

Oxidation: *loss* of one or more electrons by an element, compound, or ion.

Reduction: *gain* of one or more electrons by an element, compound, or ion.



Losing an electron means becoming less negative (i.e. more +)

Gaining an electron means becoming more negative (i.e. less +)

Oxidation–Reduction (Redox) Reaction:

Any process in which electrons are *transferred* from one substance to another.

Characteristics of Oxidation-Reduction Reactions

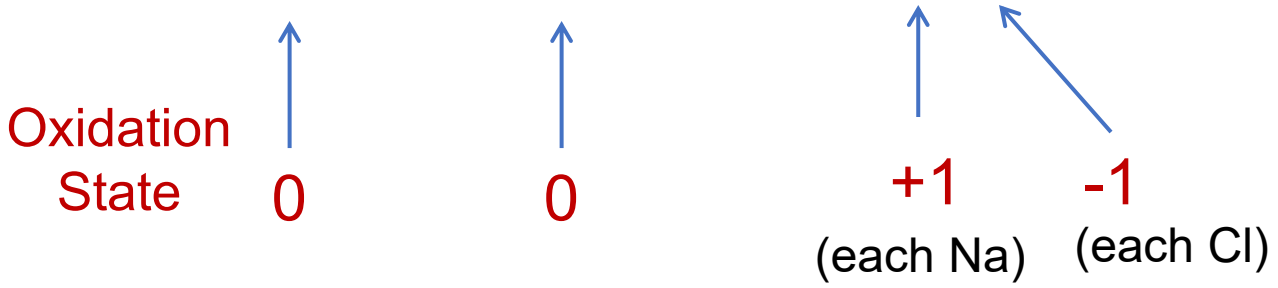
- Transfer of electrons
- Oxidation – involves an increase in oxidation state (loss of electrons) of an ***atom***; the ***compound*** is referred to as the reducing agent.
- Reduction – involves a decrease in oxidation state (gain of electrons) of an ***atom***; the ***compound*** is referred to as the oxidizing agent.

Formation of Ionic Compounds by Redox Reactions

Tracking electron transfers in the formation of ionic compounds

(Remember, sum of oxidation states of compounds is **zero**).

$$2Na(s) + Cl_2(g) \rightarrow 2NaCl(s)$$

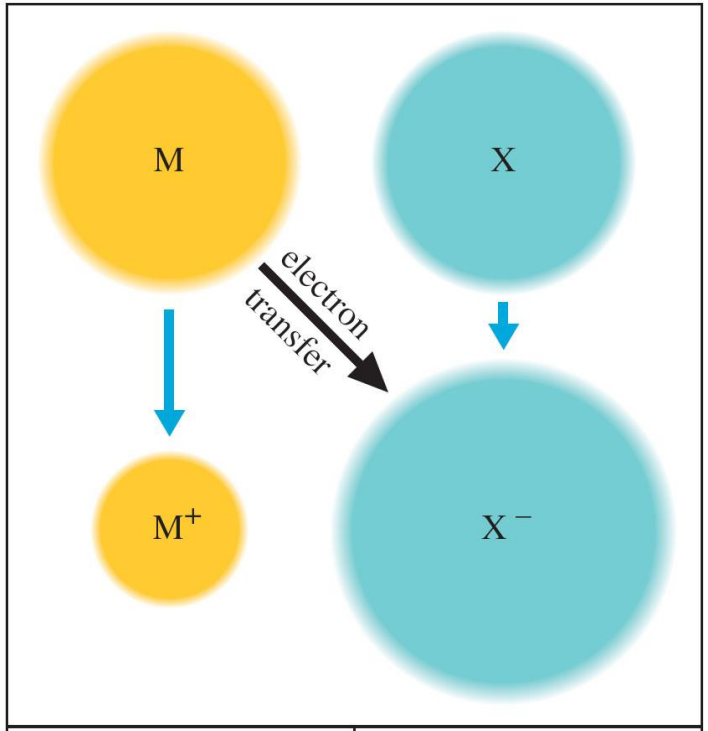


What happens?

To Na? Loses electrons: $2Na(s) \rightarrow 2Na^+(aq) + 2e^-$

To Cl? Gains electrons: $Cl_2(g) + 2e^- \rightarrow 2Cl^-(aq)$

Characteristics of Oxidation-Reduction Reactions



Oxidized	Reduced
loses electrons	*gains* electrons
oxidation state *increases*	oxidation state *decreases*
reducing agent	*oxidizing* agent

$2Na(s) + Cl_2(g) \rightarrow 2NaCl(s)$

- $2Na(s) \rightarrow 2Na^+(aq) + 2e^-$
  - Sodium atom is oxidized
  - Oxidation state increases from 0 to +1
  - Sodium metal is the reducing agent
- $Cl_2 + 2e^- \rightarrow 2Cl^-(g)$
  - Chlorine atom is reduced
  - Oxidation state decreases from 0 to -1
  - Chlorine gas is the oxidizing agent

The Oxidizing and Reducing Agents

The oxidizing agent is the **compound** containing the atoms that accept electrons

It induces something else to lose electrons.

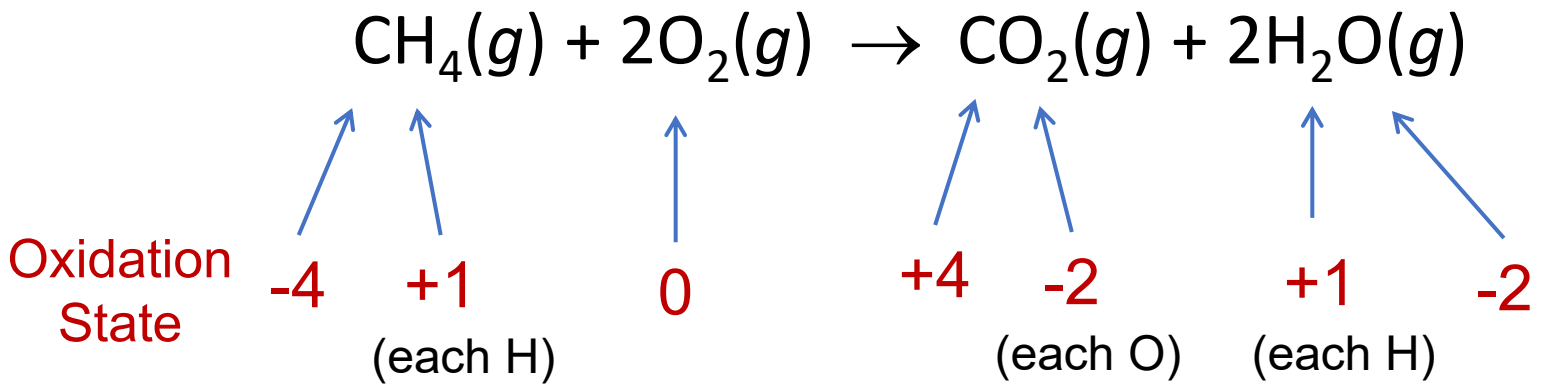
The reducing agent is the **compound** containing the atoms that lose electrons.

It induces something else to gain electrons.

Characteristics of Oxidation-Reduction Reactions

Tracking electron transfers in covalent compounds

(Remember, sum of oxidation states of compounds is **zero**).

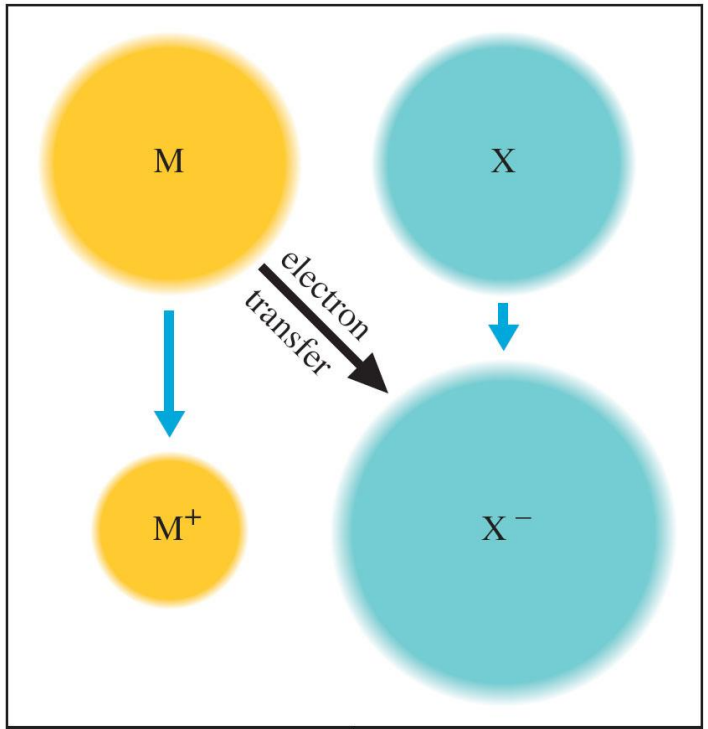


What happens?

To C? Loses electrons: $CH_4 \rightarrow CO_2 + 8e^-$

To O? Gains electrons: $2O_2 + 8e^- \rightarrow CO_2 + 2H_2O$

Characteristics of Oxidation-Reduction Reactions



Oxidized	Reduced
loses electrons	*gains* electrons
oxidation state *increases*	oxidation state *decreases*
reducing agent	*oxidizing* agent

$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$

- $CH_4 \rightarrow CO_2 + 8e^-$
  - C atom (from CH_4) is oxidized
  - Oxidation state of C increases from -4 to +4.
  - CH_4 is the reducing agent (not just the C atom)
- $2O_2 + 8e^- \rightarrow CO_2 + 2H_2O$
  - O atom is reduced
  - Oxidation state of O decreases from 0 to -2. A total of 8 electrons are gained between the two O from CO_2 and two O from H_2O.
  - O_2 is the oxidizing agent (not just the O atom)

Identifying Redox Reactions

Practice: Identify the oxidized and reduced species, along with the reducing and oxidizing agents, in this reaction.

$$4Fe + 3O_2 \rightarrow 2\ Fe_2O_3$$

Change in oxidation number:

Complete these statements using these terms:
oxidized/reduced, increased/decreased; oxidizing/reducing.

Fe: _____ Fe was _____ because its oxidation number _____.
It was the _____ agent.

O_2: _____ Oxygen was _____ because its oxidation number
_____. It was the _____ agent.

Identifying Redox Reactions

Practice: Identify the oxidized and reduced species, along with the reducing and oxidizing agents, in this reaction.

$$4Fe + 3O_2 \rightarrow 2\ Fe_2O_3$$

Change in oxidation number: 0 0 → $2(x) + 3(-2) = 0$

$x = +3$

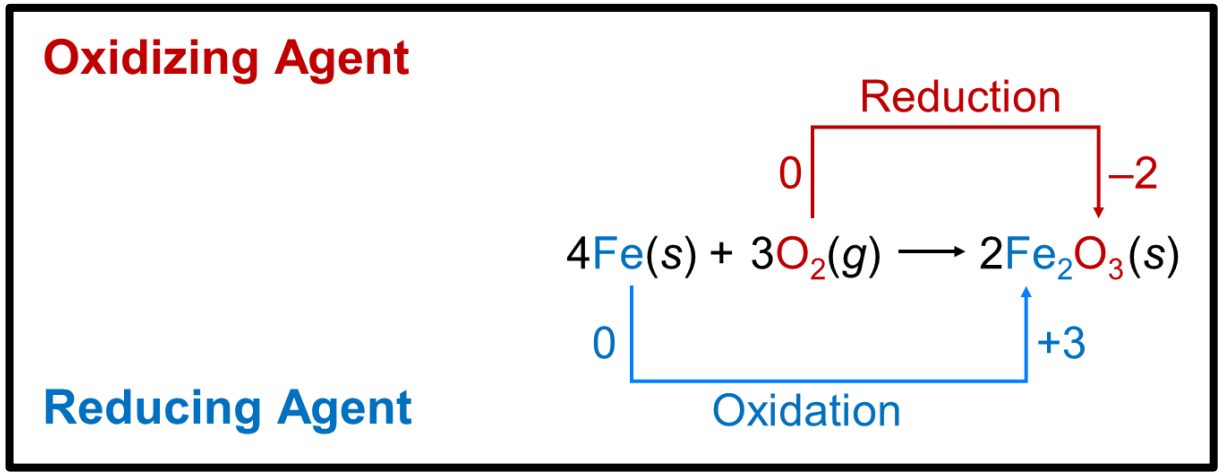
Complete these statements using these terms:
oxidized/reduced, increased/decreased; oxidizing/reducing.

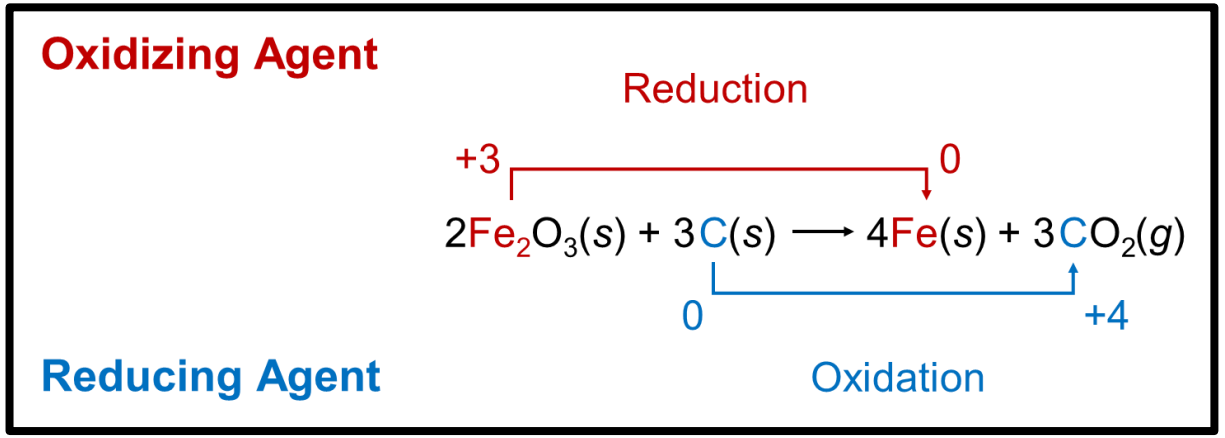
Fe: 0 to +3 Fe was oxidized because its oxidation number increased.
It was the reducing agent.

O_2: 0 to -2 Oxygen was reduced because its oxidation number decreased. It was the oxidizing agent.

Identifying Redox Reactions

Examples:





Exercise

Which of the following are oxidation-reduction reactions? Identify the oxidizing agent and the reducing agent.

a) $\mathrm{Zn_{(s)} + 2HCl_{(aq)} \rightarrow ZnCl_{2\,(aq)} + H_{2\,(g)}}$

b) $\mathrm{{Cr_2O_7}^{2-}{}_{(aq)} + 2OH^-{}_{(aq)} \rightarrow 2{CrO_4}^{2-}{}_{(aq)} + H_2O_{(l)}}$

c) $\mathrm{2CuCl_{(aq)} \rightarrow CuCl_{2\,(aq)} + Cu_{(s)}}$

Exercise

Which of the following are oxidation-reduction reactions? Identify the oxidizing agent and the reducing agent.

a) $Zn_{(s)} + 2HCl_{(aq)} \rightarrow ZnCl_{2\,(aq)} + H_{2\,(g)}$

Zn – reducing agent; HCl – oxidizing agent

b) ${Cr_2O_7}^{2-}{}_{(aq)} + 2{OH}^{-}{}_{(aq)} \rightarrow 2{CrO_4}^{2-}{}_{(aq)} + H_2O_{(l)}$

Not a redox reaction

c) $2CuCl_{(aq)} \rightarrow CuCl_{2\,(aq)} + Cu_{(s)}$

CuCl acts as the reducing and oxidizing agent

Exercises Oxidation States

1. What is the oxidation state of sulfur in S_2O_3?
2. What is the oxidation state of chromium in Na_2CrO_4?
3. What is the oxidation state of sodium in Na_2O?

Exercise: The nickel-cadmium (nicad) battery, a rechargeable "dry cell" used in battery-operated devices, uses the following redox reaction to generate electricity:

$$Cd(s) + NiO_2(s) + 2\ H_2O(l) \rightarrow Cd(OH)_2(s) + Ni(OH)_2(s)$$

Identify the substances that are oxidized and reduced, and indicate which is the oxidizing agent and which is the reducing agent.

Exercises Oxidation States

1. What is the oxidation state of sulfur in S_2O_3? +3
2. What is the oxidation state of chromium in Na_2CrO_4? +6
3. What is the oxidation state of sodium in Na_2O? +1

Exercise: The nickel-cadmium (nicad) battery, a rechargeable "dry cell" used in battery-operated devices, uses the following redox reaction to generate electricity:

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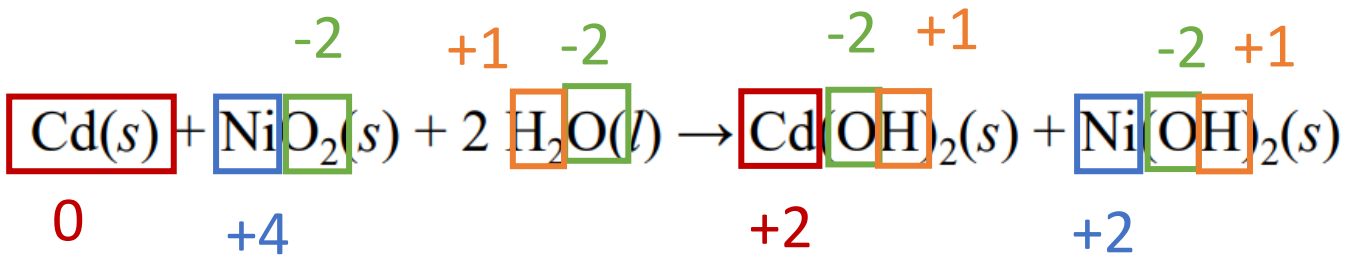
Exercises Oxidation States



Solution:

Cd: 0 to 2+

Gain/Loss e-?

Is reduced or oxidized?

Oxidizing agent or reducing agent?

Ni: +4 to +2

Gain/Loss e-?

Is reduced or oxidized?

Oxidizing agent or reducing agent?

O: -2 to -2

H: +1 to +1