Oxidation-Reduction

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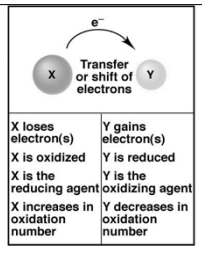
Balancing Oxidation-Reduction Reactions

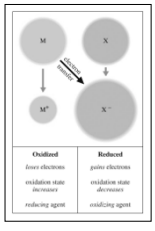
Oxidation-Reduction

- ***Oxidation*** *is the loss of electrons.*
- ***Reduction*** *is the gain of electrons.*
- *The reactions occur together. One does not occur without the other.*
- *The terms are used relative to the change in the* ***oxidation state*** *or* ***oxidation number*** *of the reactant(s).*

A Summary of Terminology for Oxidation-Reduction (Redox) Reactions

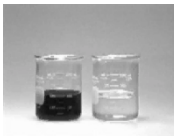


X loses electron(s)	Y gains electron(s)
X is oxidized	Y is reduced
X is the reducing agent	Y is the oxidizing agent
X increases in oxidation number	Y decreases in oxidation number



Aqueous Reactions: Oxidation - Reduction

- *In the following reaction, identify what is being oxidized and what is being reduced. What is the total number of electrons involved in the process?*



Oxidation Reduction Reactions

Oxidation-Reduction Reactions - Part 1

Oxidation-Reduction Reactions - Part 2

QUESTION

In a redox reaction, oxidation and reduction must both occur. Which statement provides an accurate premise of redox chemistry?

A. The substance that is oxidized must be the oxidizing agent.
B. The substance that is oxidized must gain electrons.
C. The substance that is oxidized must have a higher oxidation number afterwards.
D. The substance that is oxidized must combine with oxygen.

Rules for Assigning an Oxidation Number (O.N.)

General rules

1. For an atom in its elemental form (Na, O_2, Cl_2, etc.): O.N. = 0
2. For a monatomic ion: O.N. = ion charge
3. The sum of O.N. values for the atoms in a compound equals zero. The sum of O.N. values for the atoms in a polyatomic ion equals the ion charge.

Rules for specific atoms or periodic table groups

1. For Group 1A(1): O.N. = +1 in all compounds
2. For Group 2A(2): O.N. = +2 in all compounds
3. For hydrogen: O.N. = +1 in combination with nonmetals
   O.N. = −1 in combination with metals and boron
4. For fluorine: O.N. = −1 in all compounds
5. For oxygen: O.N. = −1 in peroxides
   O.N. = −2 in all other compounds (except with F)
6. For Group 7A(17): O.N. = −1 in combination with metals, nonmetals, (except O), and other halogens lower in the group

Highest and Lowest Oxidation Numbers of Reactive Main-Group Elements

Periodic Trends: Common Oxidation States

		1A	2A	3A	4A	5A	6A	7A
		+1	+2	+3	+4 / −4	+5 / −3	+6 / −2	+7 / −1
1	H	+1 / −1						
2		Li	Be	B	C	N	O	F
3		Na	Mg	Al	Si	P	S	Cl
4		K	Ca	Ga	Ge	As	Se	Br
5		Rb	Sr	In	Sn	Sb	Te	I
6		Cs	Ba	Tl	Pb	Bi	Po	At
7		Fr	Ra					

Group number
Highest O.N./Lowest O.N.

QUESTION

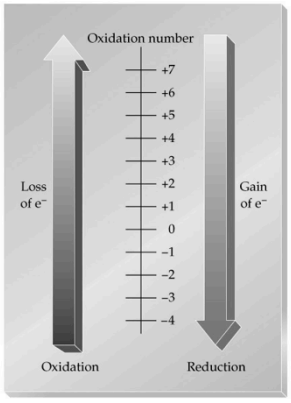
In which of the following does nitrogen have an oxidation state of +4?

A) HNO_3
B) NO_2
C) N_2O
D) NH_4Cl
E) $NaNO_2$

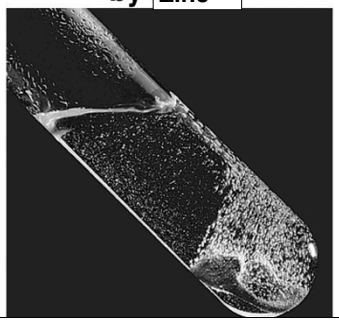
QUESTION

What is the oxidation number of chromium in ammonium dichromate?

A) +3 **B) +4** **C) +5** **D) +6**

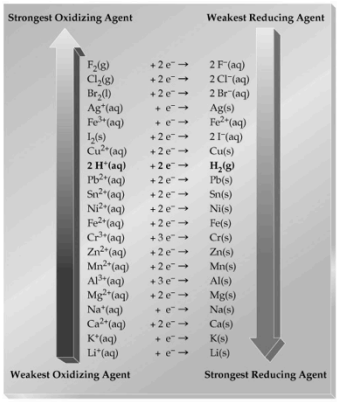


The Displacement of H_2 from Acid by Zinc



Reactivity Tables (usually reducing) show relative reactivities:
In the examples from the previous slide, the acid solution (H^+) will react with anything below it in the Table but not above.

Nickel and Zinc,
….but not Copper.



QUESTION

Select all redox reactions by looking for a change in oxidation number as reactants are converted to products.

I) $Ca + 2\,H_2O \rightarrow Ca(OH)_2 + H_2$

II) $CaO + H_2O \rightarrow Ca(OH)_2$

III) $Ca(OH)_2 + H_3PO_4 \rightarrow Ca_3(PO_4)_2 + H_2O$

IV) $Cl_2 + 2\,KBr \rightarrow Br_2 + 2\,KCl$

A) I and II **B) II and III** **C) I and IV** **D) III and IV**

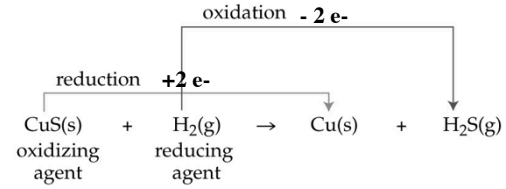
QUESTION

How many of the following are oxidation-reduction reactions?

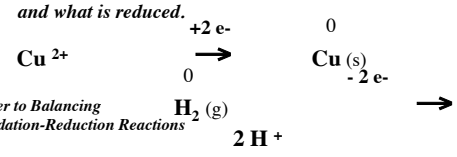
$NaOH + HCl \rightarrow NaCl + H_2O$
$Cu + 2AgNO_3 \rightarrow 2Ag + Cu(NO_3)_2$
$Mg(OH)_2 \rightarrow MgO + H_2O$
$N_2 + 3H_2 \rightarrow 2NH_3$

A) 0
B) 1
C) 2
D) 3
E) 4

Number of electrons gained must equal the number of electrons lost.



Use oxidation numbers to determine what is oxidized and what is reduced.



Refer to Balancing Oxidation-Reduction Reactions

QUESTION

In the reaction $2Cs(s) + Cl_2(g) \rightarrow 2CsCl(s)$, Cl_2 is

A. the reducing agent.
B. the oxidizing agent.
C. oxidized.
D. the electron donor.
E. two of these

Balancing Redox Equations
in acidic solutions

1) Determine the oxidation numbers of atoms in both reactants and products.
2) Identify and select out those which change oxidation number ("redox" atoms) into separate "half reactions".
3) Balance the "redox" atoms and charges (electron gain and loss must equal!).
4) In acidic reactions balance oxygen with water then hydrogen from water with acid proton(s).

Balancing Redox Equations
in acidic solutions

$Fe^{+2}_{(aq)} + Cr_2O_7^{2-}{}_{(aq)} + H^+_{(aq)} \rightarrow Fe^{3+}_{(aq)} + Cr^{3+}_{(aq)} + H_2O_{(l)}$

? Cr oxidation number?

$Fe^{2+}_{(aq)} + Cr_2O_7^{2-}{}_{(aq)} + H^+_{(aq)} \rightarrow Fe^{3+}_{(aq)} + Cr^{3+}_{(aq)} + H_2O_{(l)}$

x = ? Cr ; $2x+7(-2) = -2$; $x = +6$

Balancing Redox Equations
in acidic solutions

$Fe^{2+}_{(aq)} - e^- \rightarrow Fe^{3+}_{(aq)}$

$Cr_2O_7^{2-}{}_{(aq)} + 6\,e^- \rightarrow 2\,Cr^{3+}_{(aq)}$

Cr = (6+)

$6\,(Fe^{2+}_{(aq)} - e^- \rightarrow Fe^{3+}_{(aq)})$

$6\,Fe^{2+}_{(aq)} \rightarrow 6\,Fe^{3+}_{(aq)} + 6\,e^-$

$Cr_2O_7^{2-}{}_{(aq)} + 6\,e^- \rightarrow 2\,Cr^{3+}_{(aq)}$

Balancing Redox Equations
in acidic solutions

$6\ Fe^{2+}_{(aq)} \rightarrow 6\ Fe^{3+}_{(aq)} + 6\ e^-$

$Cr_2O_7^{2-}{}_{(aq)} + 6\ e^- \rightarrow 2\ Cr^{3+}_{(aq)}$

$6\ Fe^{2+}_{(aq)} + Cr_2O_7^{2-}{}_{(aq)} + ?\ ^{2nd}\ H^+_{(aq)} \rightarrow$
$6\ Fe^{3+}_{(aq)} + 2\ Cr^{3+}_{(aq)} + ?\ ^{1st\ Oxygen}$
$H_2O_{(l)}$

Oxygen = 7 2nd (Hydrogen) = 14

Balancing Redox Equations
in acidic solutions

Completely Balanced Equation:

$6\ Fe^{2+}_{(aq)} + Cr_2O_7^{2-}{}_{(aq)} + 14\ H^+_{(aq)} \rightarrow$
$6\ Fe^{3+}_{(aq)} + 2\ Cr^{3+}_{(aq)} + 7\ H_2O_{(l)}$

QUESTION

Dichromate ion in <u>acidic medium</u> converts ethanol, C_2H_5OH, to CO_2 according to the unbalanced equation:

$Cr_2O_7^{2-}(aq) + C_2H_5OH(aq) \rightarrow Cr^{3+}(aq) + CO_2(g) + H_2O(l)$

The <u>coefficient for H^+</u> in the balanced equation using smallest integer coefficients is:

A) 8 **B) 10** **C) 13** **D) 16**

Balancing Redox Equations
in basic solutions

1) Determine oxidation numbers of atoms in Reactants and Products
2) Identify and select out those which change oxidation number into separate "half reactions"
3) Balance redox atoms and charges (electron gain and loss must equal!)
4) In basic reactions balance the Oxygen with hydroxide then Hydrogen from hydroxide with water

Balancing Redox Equations
in basic solutions

$MnO_{2\,(aq)} + ClO_3^{1-}{}_{(aq)} + OH^{1-}{}_{aq} \rightarrow$
$MnO_4^{1-}{}_{(aq)} + Cl^{1-}{}_{(aq)} + H_2O_{(l)}$

$Mn^{4+}\,(MnO_2) \rightarrow Mn^{7+}\,(MnO_4)^{1-}$
$Cl^{+5}\,(ClO_3)^{1-} + 6\,e^- \rightarrow Cl^{1-}$

Balancing Redox Equations
in basic solutions

Electronically Balanced Equation:

$2\,MnO_{2\,(aq)} + ClO_3^{1-}{}_{(aq)} + 6\,e^- \rightarrow$
$2\,MnO_4^{1-} + Cl^{1-} + 6\,e^-$

Balancing Redox Equations
in basic solutions

Completely Balanced Equation:

$2\,MnO_{2\,(aq)} + ClO_3^{1-}{}_{(aq)} + 2\,OH^{1-}{}_{(aq)} \rightarrow$
$2\,MnO_4{}_{(aq)}^{1-} + Cl^{1-}{}_{(aq)} + 1\,H_2O_{(l)}$

9 O in product

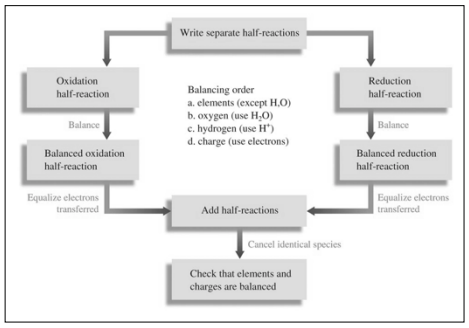
QUESTION

Oxalate ion can be found in rhubarb and spinach (among other green leafy plants). The following unbalanced equation carried out in a basic solution, shows how MnO_4^- could be used to analyze samples for oxalate.

$$MnO_4^- + C_2O_4^{2-} \rightarrow MnO_2 + CO_3^{2-} \quad \text{(basic solution)}$$

When properly balanced, how many OH^- are present?

A.1
B.2
C.3
D.4


Write separate half-reactions
Oxidation half-reaction
Balance
Balanced oxidation half-reaction
Equalize electrons transferred
Balancing order
a. elements (except H,O)
b. oxygen (use H2O)
c. hydrogen (use H+)
d. charge (use electrons)
Reduction half-reaction
Balance
Balanced reduction half-reaction
Equalize electrons transferred
Add half-reactions
Cancel identical species
Check that elements and charges are balanced