Totally Epic AP Chem Review: Oxidation Numbers and Redox Balancing!

Oxidation-Reduction (Redox) Review

Oxidation–reduction (redox) reactions: where electrons are transferred from one atom to another.

- If a substance accepts an electron, it is reduced.
- If a substance loses an electron, it is oxidized.

Two great mnemonics!

1. OIL RIG: **O**xidation **I**s **L**oss (OIL) and **R**eduction **I**s **G**ain (RIG)
2. LEO goes GER: A species **loses electrons** when **oxidized**, and **gains electrons** when **reduced.**

If a chemical reaction has two species which change oxidation number: yes, **it's redox!**

Oxidation Numbers/States: Oxidation states are imaginary charges assigned based on a set of rules simply used to determine electron flow. Even though they look like them, oxidation states are **NOT** ionic charges!

Rules for Assigning Oxidation Numbers (in order of priority)

If any rules are in conflict, follow the rule that is higher on the list!

Oxidation Rules:

1. Free elements = 0
2. All atoms in a neutral compound add up to 0.
3. All atoms in a polyatomic ion add up to the ion's charge.

The rules below apply to bonded elements:

4. Group 1A metals = +1
5. Group 2A = +2
6. Non-metals **usually** follow the chart to the right, in order:
   a. Note: the carbon family (4A) isn't mentioned – you will ALWAYS have to solve for the oxidation number of group 4A elements in a compound.

Nonmetal	Oxidation State	Example
Fluorine	−1	MgF_2 −1 ox state
Hydrogen	+1	H_2O +1 ox state
Oxygen	−2	CO_2 −2 ox state
Group 7A	−1	CCl_4 −1 ox state
Group 6A	−2	H_2S −2 ox state
Group 5A	−3	NH_3 −3 ox state

Oxidation State Practice: Give the oxidation number of each element for the chemical species listed below.

1. $SO_3^{2-}(aq)$ +4, -2
2. $KMnO_4(aq)$ +1, +7, -2
3. $H_2O_2(l)$ +1, -1
4. $Fe(ClO_4)_3(s)$ +3, +7, -2
5. $NaH(s)$ +1, -1
6. $C_2H_4O_2(aq)$ Ø, +1, -2

How to Identify What is Oxidized or Reduced in a Reaction

A substance that has the element that has been oxidized (LOST electrons) will have an oxidation number that becomes more positive (or less negative).

A substance that has the element that has been reduced (GAINED electrons) will have an oxidation number that becomes more negative (or less positive).

$$Zn(s) + Fe^{2+}(aq) \rightarrow Zn^{2+}(aq) + Fe(s)$$

Oxidation states: $\emptyset$ $+2$ $+2$ $\emptyset$

What was oxidized? $Zn(s)$ What was reduced? $Fe^{2+}(aq)$

$$2\,C_2H_2(g) + 5\,O_2(g) \rightarrow 4\,CO_2(g) + 2\,H_2O(g)$$

Oxidation states: $-1, +1$ $\emptyset$ $+4, -2$ $+1, -2$

What was oxidized? C in $C_2H_2(g)$ What was reduced? $O_2(g)$

Steps to Balance Redox Reactions using the Half-Reaction Method

1. Write the oxidation and reduction half-reactions (without electrons).
2. Balance all atoms other than H and O.
3. Balance O by adding H_2O, and balance H by adding H^+.
4. Balance charge in each half reaction using electrons.
5. Equalize electron transfer by multiplying so both half-reactions have the same # of electrons exchanged.
6. Add the two half-reactions, cancelling anything that appears on both sides of the equation.
7. **In a basic solution:** add OH^- ions to ~~each~~ equal/both side of the equation to neutralize H^+ ions. Combine H^+ and OH^- to form H_2O, and cancel if needed.

Practice:

1. Balance the following redox reaction: $2\,Ag^+(aq) + 1\,Cu(s) \rightarrow 2\,Ag(s) + 1\,Cu^{2+}(aq)$

$2\,(Ag^+ + 1e^- \rightarrow Ag)$ $\quad Cu \rightarrow Cu^{2+} + 2e^-$

2. Balance the following redox reaction: $14\,H^+ + 1\,Cr_2O_7^{2-} + 3\,S^{2-} \rightarrow 2\,Cr^{3+} + 3\,S + 7\,H_2O$

$Cr_2O_7^{2-} \rightarrow 2Cr^{3+}$ $\quad (S^{2-} \rightarrow S + 2e^-) \times 3$

$14H^+ + Cr_2O_7^{2-} \rightarrow 2Cr^{3+} + 7H_2O$ $\quad 3S^{2-} \rightarrow 3S + 6e^-$

$12+ \qquad 6+$

$14H^+ + Cr_2O_7^{2-} + 6e^- \rightarrow 2Cr^{3+} + 7H_2O$

3. Balance the following redox reaction: $\underline{4}\ H_2O + \underline{2}\ MnO_4^- + \underline{6}\ I^- \rightarrow \underline{3}\ I_2 + \underline{2}\ MnO_2 + \underline{8}\ OH^-$

$MnO_4^- \rightarrow MnO_2$

$4H^+ + MnO_4^- \rightarrow MnO_2 + 2H_2O$

(+3 → 0)

$(4H^+ + MnO_4^- + 3e^- \rightarrow MnO_2 + 2H_2O) \times 2$

$8H^+ + 2MnO_4^- + 6e^- \rightarrow 2MnO_2 + 4H_2O$

$2I^- \rightarrow I_2$

$(2I^- \rightarrow I_2 + 2e^-) \times 3$

$6I^- \rightarrow 3I_2 + 6e^-$

+

Add same # to both sides

$8H^+ + 2MnO_4^- + 6I^- \rightarrow 3I_2 + 2MnO_2 + 4H_2O + 8OH^-$

basic sol'n!

$+ 8OH^-$

$= 8H_2O$

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$$2\ H_2O(l) + 4\ MnO_4^-(aq) + 3\ ClO_2^-(aq) \rightarrow 4\ MnO_2(aq) + 3\ ClO_4^-(aq) + 4\ OH^-(aq)$$

+1,-2 +7,-2 +3,-2 +4,-2 +7,-2 -2,+1

4. Which species is reduced in the reaction represented above?

a. MnO_2 b. ClO_2^- (c.) MnO_4^- d. ClO_4^-

5. In the reaction below, a piece of solid nickel is added to a solution of potassium dichromate.

$$14\ H^+(aq) + Cr_2O_7^{2-}(aq) + 3\ Ni(s) \rightarrow 2\ Cr^{3+}(aq) + 3\ Ni^{2+}(aq) + 7\ H_2O(l)$$

+1 +6,-2 0 +3 +2 +1,-2

Which species is being oxidized and which is being reduced?

	Oxidized	Reduced
a.	$Cr_2O_7^{2-}(aq)$	$Ni(s)$
b.	$Cr^{3+}(aq)$	$Ni^{2+}(aq)$
(c.)	$Ni(s)$	$Cr_2O_7^{2-}(aq)$
d.	$Ni^{2+}(aq)$	$Cr^{3+}(aq)$

6. Which net ionic equation below represents a possible reaction that takes place when a strip of magnesium metal is oxidized by a solution of chromium (III) nitrate?

a. $Mg(s) + Cr(NO_3)_3(aq) \rightarrow Mg^{2+}(aq) + Cr^{3+}(aq) + 3\ NO_3^-(aq)$

(b.) $3\ Mg(s) + 2\ Cr^{3+}(aq) \rightarrow 3\ Mg^{2+}(aq) + 2\ Cr(s)$ (net ionic = no NO_3^-!)

c. $Mg(s) + Cr^{3+}(aq) \rightarrow Mg^{2+}(aq) + Cr(s)$

d. $3\ Mg(s) + 2\ Cr(NO_3)_3(aq) \rightarrow 3\ Mg^{2+}(aq) + 2\ Cr(s) + NO_3^-(aq)$