SAMPLE DELIVERY: NOTES CLEANED INTO A REVISION PACK

Redox Reactions Mastery Guide

Use this as a study note, project reference, or revision sheet. Redox becomes easy once you can track electrons and oxidation numbers without panicking.

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1. The Core Idea

`Redox` means `reduction + oxidation`.

A redox reaction is a chemical reaction where electrons are transferred, or where the oxidation numbers of atoms change.

Classic example:

```
Zn(s) + Cu2+(aq) -> Zn2+(aq) + Cu(s)
```

What happens:

```
Zn loses 2 electrons:
Zn(s) -> Zn2+(aq) + 2e-

Cu2+ gains 2 electrons:
Cu2+(aq) + 2e- -> Cu(s)
```

Overall:

```
Zn(s) + Cu2+(aq) -> Zn2+(aq) + Cu(s)
```

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2. Essential Definitions

Term	Meaning	Electron view	Oxidation-number view
Oxidation	Loss of electrons	Electrons are produced	Oxidation number increases
Reduction	Gain of electrons	Electrons are consumed	Oxidation number decreases
Oxidizing agent	Substance that causes oxidation	Gains electrons	Gets reduced
Reducing agent	Substance that causes reduction	Loses electrons	Gets oxidized
Half-reaction	One side of a redox process	Shows either oxidation or reduction	Used for balancing

Memory tools:

```
OIL RIG: Oxidation Is Loss, Reduction Is Gain
LEO GER: Loss of Electrons is Oxidation, Gain of Electrons is Reduction
An Ox, Red Cat: Oxidation at Anode, Reduction at Cathode
```

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3. Electron Transfer Diagram

```
Reducing agent                    Oxidizing agent
gets oxidized                     gets reduced

Zn(s)                             Cu2+(aq)
```

```
oxidation number 0                    oxidation number +2

        2 electrons move from Zn to Cu2+
Zn  ------------------------------------>  Cu2+

Zn2+(aq)                               Cu(s)
oxidation number +2                    oxidation number 0
```

4. Concept Map

```
flowchart LR
    A[Reducing agent] -->|loses electrons| B[Oxidized]
    B -->|oxidation number increases| C[Product with higher oxidation state]
    D[Oxidizing agent] -->|gains electrons| E[Reduced]
    E -->|oxidation number decreases| F[Product with lower oxidation state]
    A -. causes .-> E
    D -. causes .-> B
```

5. Oxidation Number Rules

These rules are your main tool for identifying redox reactions.

Rule	Example
Free elements have oxidation number 0.	`Na`, `O2`, `Cl2`, `Fe` are 0
A monatomic ion has oxidation number equal to its charge.	`Na+` = +1, `Cl-` = -1, `Fe3+` = +3
Group 1 metals are usually +1.	`Na`, `K`, `Li`
Group 2 metals are usually +2.	`Mg`, `Ca`, `Ba`
Fluorine is almost always -1.	`HF`, `NaF`
Oxygen is usually -2.	`H2O`, `CO2`, `SO4^2-`
Oxygen is -1 in peroxides.	`H2O2`
Hydrogen is usually +1 with nonmetals.	`H2O`, `HCl`, `CH4`
Hydrogen is -1 in metal hydrides.	`NaH`, `CaH2`
Sum of oxidation numbers in a neutral compound is 0.	`H2O`: 2(+1) + (-2) = 0
Sum of oxidation numbers in a polyatomic ion equals the ion charge.	`SO4^2-`: S + 4(-2) = -2, so S = +6

6. How To Identify A Redox Reaction

1. Assign oxidation numbers to the important atoms.
2. Compare reactants and products.
3. If any atom increases in oxidation number, it is oxidized.
4. If any atom decreases in oxidation number, it is reduced.
5. If both happen, the reaction is redox.

Example:

```
Zn(s) + Cu2+(aq) -> Zn2+(aq) + Cu(s)

Zn: 0 -> +2  oxidized
Cu: +2 -> 0  reduced
```

So:

```
Zn is the reducing agent because it is oxidized.
Cu2+ is the oxidizing agent because it is reduced.
```

---

7. Common Redox Reaction Types

Single Displacement

```
Zn(s) + CuSO4(aq) -> ZnSO4(aq) + Cu(s)
```

Net ionic equation:

```
Zn(s) + Cu2+(aq) -> Zn2+(aq) + Cu(s)
```

Zinc displaces copper because zinc more easily loses electrons.

Combustion

```
CH4 + 2O2 -> CO2 + 2H2O
```

Carbon in methane:

```
C in CH4: -4
C in CO2: +4
```

Carbon is oxidized. Oxygen goes from `0` in `O2` to `-2` in products, so oxygen is reduced.

Corrosion

Rusting involves iron being oxidized and oxygen being reduced.

Simplified idea:

```
Fe -> Fe2+ + 2e-
O2 gains electrons in the presence of water
```

Disproportionation

One element is both oxidized and reduced.

```
Cl2 + 2OH- -> Cl- + ClO- + H2O
```

Chlorine:

```
Cl in Cl2: 0
Cl in Cl-: -1  reduced
Cl in ClO-: +1 oxidized
```

Comproportionation

Two oxidation states of the same element form one middle oxidation state.

```
Fe(s) + 2Fe3+(aq) -> 3Fe2+(aq)
```

Iron goes from `0` and `+3` to `+2`.

8. Balancing Redox Reactions: Acidic Solution

Use the half-reaction method.

Steps:

1. Split the reaction into oxidation and reduction half-reactions.
2. Balance all atoms except `O` and `H`.
3. Balance `O` using `H2O`.
4. Balance `H` using `H+`.
5. Balance charge using electrons, `e-`.
6. Multiply half-reactions so electrons lost = electrons gained.
7. Add the half-reactions.
8. Cancel species appearing on both sides.
9. Check atoms and total charge.

Worked Example: Permanganate and Iron(II) in Acid

Unbalanced:

```
MnO4- + Fe2+ -> Mn2+ + Fe3+
```

Reduction half-reaction:

```
MnO4- -> Mn2+
```

Balance oxygen using water:

```
MnO4- -> Mn2+ + 4H2O
```

Balance hydrogen using `H+`:

```
MnO4- + 8H+ -> Mn2+ + 4H2O
```

Balance charge using electrons:

```
MnO4- + 8H+ + 5e- -> Mn2+ + 4H2O
```

Oxidation half-reaction:

```
Fe2+ -> Fe3+ + e-
```

Multiply iron half-reaction by 5:

```
5Fe2+ -> 5Fe3+ + 5e-
```

Add and cancel electrons:

```
MnO4- + 8H+ + 5Fe2+ -> Mn2+ + 4H2O + 5Fe3+
```

Check charge:

```
Left: -1 + 8(+1) + 5(+2) = +17
Right: +2 + 5(+3) = +17
```

Balanced.

9. Balancing Redox Reactions: Basic Solution

Basic solution uses `OH-`. The easiest method:

1. Balance as if the solution were acidic.
2. Add `OH-` to both sides to neutralize every `H+`.
3. Combine `H+ + OH-` into `H2O`.
4. Cancel extra water molecules.
5. Check atoms and charge.

Worked Example: Permanganate and Iodide in Basic Solution

Unbalanced:

```
MnO4- + I- -> MnO2 + I2
```

Reduction in basic solution:

```
MnO4- + 2H2O + 3e- -> MnO2 + 4OH-
```

Oxidation:

```
2I- -> I2 + 2e-
```

Equalize electrons:

```
2(MnO4- + 2H2O + 3e- -> MnO2 + 4OH-)
3(2I- -> I2 + 2e-)
```

Add:

```
2MnO4- + 4H2O + 6I- -> 2MnO2 + 8OH- + 3I2
```

Check charge:

```
Left: 2(-1) + 6(-1) = -8
Right: 8(-1) = -8
```

Balanced.

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10. Electrochemistry Connection

Redox reactions are the basis of batteries and electrochemical cells.

For a zinc-copper galvanic cell:

```
Anode:   Zn(s) -> Zn2+(aq) + 2e-       oxidation
Cathode: Cu2+(aq) + 2e- -> Cu(s)       reduction
Overall: Zn(s) + Cu2+(aq) -> Zn2+(aq) + Cu(s)
```

Cell notation:

```
Zn(s) | Zn2+(aq) || Cu2+(aq) | Cu(s)
```

Meaning:

```
single line |  = phase boundary
double line || = salt bridge
left side     = anode
right side    = cathode
```

In a galvanic cell:

```
Electrons flow through the wire from anode to cathode.
Oxidation happens at the anode.
Reduction happens at the cathode.
The salt bridge keeps charge balanced.
```

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11. Common Oxidizing and Reducing Agents

Common Oxidizing Agents

Oxidizing agents get reduced.

```
O2
Cl2, Br2, I2
KMnO4 / MnO4-
K2Cr2O7 / Cr2O7^2-
H2O2
HNO3
Cu2+
Fe3+
ClO-
```

Common Reducing Agents

Reducing agents get oxidized.

```
Reactive metals: Zn, Mg, Fe, Al
H2
C
CO
I-
Fe2+
SO2
H2S
S2O3^2-
```

Important note: whether something acts as an oxidizing or reducing agent depends on the reaction conditions.

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12. Fast Exam Method

When you see a redox question:

```
1. Assign oxidation numbers.
2. Mark increase as oxidation.
3. Mark decrease as reduction.
4. Name agents:
   oxidized species = reducing agent
   reduced species = oxidizing agent
5. If balancing:
   split half-reactions
   balance atoms
```

```
balance charge with electrons
equalize electrons
add and cancel
```

---

13. Common Mistakes

Mistake	Fix
Saying the oxidizing agent is oxidized	Oxidizing agent gets reduced
Putting electrons on the wrong side	Oxidation produces electrons; reduction consumes electrons
Forgetting to balance charge	Always check total charge at the end
Balancing atoms only	Redox equations must balance atoms and charge
Ignoring spectator ions	Use net ionic equations when possible
Forgetting basic solution cleanup	Add `OH-` to neutralize `H+`, then cancel water

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14. Practice Problems

A. Oxidation Numbers

Find the oxidation number of the named element.

```
1. Mn in MnO4-
2. S in SO4^2-
3. Cl in ClO3-
4. Cr in Cr2O7^2-
5. N in NH4+
```

Answers:

```
1. Mn = +7
2. S = +6
3. Cl = +5
4. Cr = +6
5. N = -3
```

B. Identify Oxidation, Reduction, and Agents

Problem 1:

```
Zn + 2H+ -> Zn2+ + H2
```

Answer:

```
Zn: 0 -> +2, oxidized, reducing agent
H: +1 -> 0, reduced, oxidizing agent
```

Problem 2:

```
Cl2 + 2Br- -> 2Cl- + Br2
```

Answer:

```
Br-: -1 -> 0, oxidized, reducing agent
Cl2: 0 -> -1, reduced, oxidizing agent
```

C. Balance These Redox Equations

Problem 1, acidic:

```
Cr2O7^2- + Fe2+ -> Cr3+ + Fe3+
```

Answer:

```
Cr2O7^2- + 14H+ + 6Fe2+ -> 2Cr3+ + 7H2O + 6Fe3+
```

Problem 2, basic:

```
ClO- + I- -> Cl- + I2
```

Answer:

```
ClO- + H2O + 2I- -> Cl- + 2OH- + I2
```

Problem 3, disproportionation:

```
H2O2 -> O2 + H2O
```

Answer:

```
2H2O2 -> O2 + 2H2O
```

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15. Mastery Checklist

You understand redox when you can do these without notes:

```
[ ] Define oxidation and reduction using electrons.
[ ] Define oxidation and reduction using oxidation numbers.
[ ] Assign oxidation numbers in compounds and ions.
[ ] Identify oxidized and reduced species.
[ ] Identify oxidizing and reducing agents.
[ ] Split a reaction into half-reactions.
[ ] Balance redox reactions in acidic solution.
[ ] Balance redox reactions in basic solution.
[ ] Explain a galvanic cell using anode, cathode, electrons, and salt bridge.
[ ] Recognize redox in corrosion, combustion, respiration, photosynthesis, and batteries.
```

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16. Five-Day Mastery Plan

Day 1:

```
Learn definitions and oxidation number rules.
Do 20 oxidation-number questions.
```

Day 2:

```
Identify oxidation, reduction, oxidizing agents, and reducing agents.
Do 15 reaction-identification questions.
```

Day 3:

```
Balance redox equations in acidic solution.
Focus on half-reactions.
```

Day 4:

```
Balance redox equations in basic solution.
Practice converting H+ to H2O using OH-.
```

Day 5:

```
Study electrochemical cells.
Explain anode, cathode, electron flow, salt bridge, and cell notation.
Then do mixed review.
```

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17. Project Summary Paragraph

Redox reactions are oxidation-reduction reactions in which electrons are transferred between chemical species, causing changes in oxidation numbers. Oxidation is the loss of electrons and an increase in oxidation number, while reduction is the gain of electrons and a decrease in oxidation number. The substance that loses electrons is oxidized and acts as the reducing agent, while the substance that gains electrons is reduced and acts as the oxidizing agent. Redox reactions are important in batteries, corrosion, combustion, bleaching, respiration, and photosynthesis. They can be balanced using the half-reaction method, which separately balances oxidation and reduction before combining them so that electrons cancel.

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18. Sources For Further Study

- OpenStax Chemistry, "Balancing Oxidation-Reduction Reactions": https://openstax.org/books/chemistry/pages/17-1-balancing-oxidation-reduction-reactions
- OpenStax Chemistry 2e, "Classifying Chemical Reactions": https://openstax.org/books/chemistry-2e/pages/4-2-classifying-chemical-reactions
- Chemistry LibreTexts, "Oxidation-Reduction Reactions": https://chem.libretexts.org/Bookshelves/General_Chemistry/Book%3A_General_Chemistry%3A_Principles_Patterns_and_Applications_Reduction_Reactions
- Khan Academy, "Redox reactions and electrochemistry": https://www.khanacademy.org/science/chemistry/oxidation-reduction
- Britannica, "Oxidation-reduction reaction": https://www.britannica.com/science/oxidation-reduction-reaction

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Fast Practice MCQs

1. Oxidation means: A. gain of electrons B. loss of electrons C. no electron change D. only oxygen removal
2. In `Zn -> Zn2+ + 2e-`, zinc is: A. reduced B. oxidized C. neutralized D. precipitated
3. The oxidizing agent: A. gains electrons B. loses electrons C. always releases gas D. never changes oxidation number
4. OIL RIG means: A. oxidation is loss, reduction is gain B. oxidation is light, reduction is gas C. oxygen is low, reduction is high D. only ions lose electrons
5. In a galvanic cell, oxidation occurs at the: A. cathode B. salt bridge C. anode D. electrolyte only

Answer key: 1-B, 2-B, 3-A, 4-A, 5-C

Service idea: send messy notes, PDF, or chapter photos. Receive a clean study pack with summary, MCQs, and answer key.