Unit 16: Electrochemistry

Redox Reactions

This worksheet will cover **reduction-oxidation (redox) reactions**, highlighting **oxidation states** and **reducing/oxidizing agents**. It will discuss determining oxidation numbers for ions, compounds, and free elements, as well as which species are being reduced and oxidized in a reaction. As you progress through the worksheet, you will develop the skills necessary to analyze redox reactions and oxidation states.

Practice Problems:

1. Determine the oxidation number of each atom/ion.

   a. Sn b. Fe^{+3} c. Sn^{+4} d. nitrate e. Sulfate f. O_2

2. Calculate the oxidation number of sulfur in each of the following.

   a. S b. SO_3^{-2} c. SCl_6 d. Na_2SO_4

3. Distinguish between an oxidizing agent and a reducing agent.

4. Identify which atom is oxidized and which atom is reduced. Then determine which reactant is the oxidizing agent and which is the reducing agent. (Hint: determine the oxidation numbers of each atom before and after the reaction.)

   a. $2\ Li + F_2 \rightarrow 2\ LiF$

Oxidized ____________________ Reduced ____________________

Oxidizing Agent ____________________ Reducing Agent ____________________

   b. $S + Cl_2 \rightarrow SCl_2$

Oxidized ____________________ Reduced ____________________

Oxidizing Agent ____________________ Reducing Agent ____________________

   c. $2\ O_2 + N_2 \rightarrow 2\ NO_2$

Oxidized ____________________ Reduced ____________________

Oxidizing Agent __________________ Reducing Agent __________________

d. $Mg + Cu(NO_3)_2 \rightarrow Mg(NO_3)_2 + Cu$

Oxidized __________________ Reduced __________________

Oxidizing Agent __________________ Reducing Agent __________________

5. Write half-reactions for the oxidation and reduction process for each of the following.

a. $Fe^{+2}_{(aq)} + Co_{(s)} \rightarrow Co^{+2}_{(aq)} + Fe_{(s)}$

i) Oxidation:
ii) Reduction:

b. $Sn_{(s)} + O_{2(g)} \rightarrow O^{-2} + 2\ Sn^{+2}_{(aq)}$

i) Oxidation:
ii) Reduction:

c. $Cu^{+2}_{(aq)} + Pb_{(s)} \rightarrow Cu_{(s)} + Pb^{+2}_{(aq)}$

i) Oxidation:
ii) Reduction:

ANSWER KEY

Redox Reactions

Practice Problems:

1. Determine the oxidation number of each atom/ion.

a. Sn	b. Fe^{+3}	c. Sn^{+4}	d. nitrate	e. Sulfate	f. O_2
0	+3	+4	-1	-2	0
(free element)	(ion charge)	(ion charge)	(polyatomic ion charge)	(polyatomic ion charge)	(free element)

Rules for oxidation numbers:

Oxidation numbers have certain rules when being assigned to elements in a compound.

- Oxidation number of a free element (O_2 (g), Zn (s), Cu (s), etc.) will always be 0.
- Sum of oxidation numbers in a neutral compound will add up to 0.
- Sum of oxidation numbers in a polyatomic ion = charge of that ion.
- Elements may have multiple oxidation states.
    - Transition metals usually have multiple; for example, chromium has oxidation states of +2, +3, and +6.
- Certain elements usually have specific oxidation numbers when in compounds.
    - Alkali metals = +1.
    - Alkali earth metals = +2.
    - Oxygen = -2 (except in peroxide ion (O_2^{-2}), where it is -1).
    - Fluorine = -1. (other halogens may have positive values when in polyatomic ions with oxygen).
    - Hydrogen = +1 (except when bonded with metals, where it will be -1).
    - Metals are always positive; non-metals may have positive or negative oxidation numbers.
    - In an ionic compound with 2 elements, the metal or the more electronegative atom has a positive oxidation number, and the non-metal or less electronegative atom has a negative oxidation number.
    - The charge of a monatomic ion (ion that consists of 1 type of atom) can tell you the oxidation state. For example, nitrogen is -3 in the nitride ion, N^{-3}.

a. Sn is a free element, so its oxidation number will be 0.

b. Fe^{+3} is a monatomic ion with a charge of +3, so its oxidation number will match its charge.

c. Sn^{+4} is also a monatomic ion with a charge of +4, so its oxidation number will match its charge.

d. Nitrate is the polyatomic ion NO_3^-. The question asks for the oxidation state of the entire ion, so that will match its charge of -1.

e. Sulfate is the polyatomic ion SO_4^{-2}. The question asks for the oxidation state of the entire ion, so that will match its charge of -2.

f. O_2 is a free element (since oxygen is diatomic), so the oxidation number of the entire compound will be 0.

2. Calculate the oxidation number of sulfur in each of the following.

a. S	b. SO_3^{-2}	c. SCl_6	d. Na_2SO_4
0	+4	+6	+6

a. Sulfur by itself (S) is a free element, which has oxidation states of 0.
b. This sulfite ion charge is -2, so sulfur and oxygen's oxidation states must add up to -2. Oxygen has an established oxidation state of -2, so we can set up a small equation as such.

$$S + 3\,O = -2$$
$$S + 3\,(-2) = -2$$
$$S = +4$$

c. Begin with the most electronegative atom. Since chlorine, a halogen, is not in a polyatomic ion with oxygen, we can give chlorine the oxidation state of -1 (its usual periodic table charge). Set up a small equation as such, knowing that this compound is neutral and its total oxidation number must equal 0.

$$S + 6\,Cl = 0$$
$$S + 6\,(-1) = 0$$
$$S = +6$$

d. The sum of oxidation numbers in a neutral compound, such as Na_2SO_4 equals 0. We can split up an ionic compound into its individual ions, Na^{+2} and SO_4^{-2} to determine the oxidation state of sulfur.

$$S + 4\,O = -2$$
$$S + 4\,(-2) = -2$$
$$S = +6$$

3. Distinguish between an oxidizing agent and a reducing agent.

An oxidizing agent is a chemical species that forces oxidation to occur and is reduced itself. A reducing agent is a chemical species that forces reduction to occur and is oxidized itself.

4. Identify which atom is oxidized and which atom is reduced. Then determine which reactant is the oxidizing agent and which is the reducing agent. (Hint: determine the oxidation numbers of each atom before and after the reaction.)

a. $2\,Li + F_2 \rightarrow 2\,LiF$

Oxidized	Li	Reduced	F
Oxidizing Agent	F_2	Reducing Agent	Li

First, we can determine the oxidation numbers of each atom before and after the reaction.

Atom	Before	After
Li	0 (free element)	+1 (ionic charge as Li^+)
F	0 (free element)	-1 (ionic charge as F^-)

Since the oxidation number of Li increases, Li is becoming more positive and losing electrons, which indicates Li as being oxidized. F is becoming more negative and gains electrons, indicating F is being reduced. The entire F_2 species is forcing the oxidation of Li, so F_2 is the oxidizing agent. Li is the reducing agent as it forces F to be reduced.

b. $S + Cl_2 \rightarrow 2\ SCl_2$

Oxidized ________S________ Reduced ________Cl________

Oxidizing Agent ________Cl_2________ Reducing Agent ________S________

Determine the oxidation numbers of each atom before and after the reaction.

Atom	Before	After
S	0 (free element)	+2 S + 2 Cl = 0 S + 2 (-1) = 0 S = +2
Cl	0 (free element)	-1 (halogen)

Since sulfur's oxidation state becomes more positive, sulfur is losing electrons and becomes oxidized. The oxidizing agent is therefore Cl_2. Cl is gaining electrons and becomes reduced. The reducing agent is S.

c. $2\ O_2 + N_2 \rightarrow 2\ NO_2$

Oxidized ________N________ Reduced ________O________

Oxidizing Agent ________O_2________ Reducing Agent ________N_2________

Find the oxidation states of each atom.

Atom	Before	After

O	0 (free element)	-2 (established charge)
N	0 (free element)	+4 N + 2 O = 0 N + 2 (-2) = 0 N = +4

Oxygen's oxidation number is decreasing, indicating the oxygen atom is gaining electrons and is being reduced. N_2 is forcing the reduction, meaning N_2 is the reducing agent. Nitrogen's oxidation number is increasing, indicating the nitrogen atom is losing electrons (becoming more positive) and is being oxidized. O_2 is forcing the oxidation, meaning O_2 is the oxidizing agent.

d. $Mg + Cu(NO_3)_2 \rightarrow Mg(NO_3)_2 + Cu$

Oxidized ________Mg________ Reduced ________Cu^{+2}_______

Oxidizing Agent ______$Cu(NO_3)_2$______ Reducing Agent ________Mg________

Find the oxidation states of each atom before and after the reaction.

Atom	Before	After
Mg	0 (free element)	+2 (established charge)
Cu	+2 (ionic charge in $Cu(NO_3)_2$)	0 (free element)
N	+5 in NO_3^- N + 3 O = -1 N + 3 (-2) = -1 N + -6 = -1 N = +5	+5 in NO_3^- N + 3 O = -1 N + 3 (-2) = -1 N + -6 = -1 N = +5
O	-2 (established charge)	-2 (established charge)

5. Write half-reactions for the oxidation and reduction process for each of the following.

a. $Fe^{+2}_{(aq)} + Co_{(s)} \rightarrow Co^{+2}_{(aq)} + Fe_{(s)}$

i) Oxidation: $Co \rightarrow Co^{+2} + 2e^-$
ii) Reduction: $Fe^{+2} + 2e^- \rightarrow Fe$

Fe^{+2} becomes Fe, meaning it is gaining electrons as it loses its positive charge. Therefore, we can write a reduction reaction for Fe. Co is losing electrons to become Co^{+2}, so we can write an oxidation reaction for Co.

b. $2\ Sn_{(s)} + O_{2(g)} \rightarrow O^{-2} + 2\ Sn^{+2}_{(aq)}$

i) Oxidation: $Sn \rightarrow Sn^{+2} + 2e^{-}$
ii) Reduction: $O_2 + 4e^{-} \rightarrow 2O^{-2}$

O_2 becomes O^{-2}, meaning it is gaining electrons as it loses its neutral state. Therefore, we can write a reduction reaction for O. Sn is losing electrons to become Sn^{+2}, so we can write an oxidation reaction for Sn.

c. $Cu^{+2}_{(aq)} + Pb_{(s)} \rightarrow Cu_{(s)} + Pb^{+2}_{(aq)}$

i) Oxidation: $Pb \rightarrow Pb^{+2} + 2e^{-}$
ii) Reduction: $Cu^{+2} + 2e^{-} \rightarrow Cu$

Cu^{+2} becomes Cu, meaning it is gaining electrons as it loses its positive charge. Therefore, we can write a reduction reaction for Cu. Pb is losing electrons to become $Pb^{"2}$, so we can write an oxidation reaction for Pb.