B2 · Oxidation numbers and redox changes

Completed notes

Goals: Find an oxidation number with the rules and the sum rule. Use oxidation numbers to find what is oxidized and what is reduced. Name both agents, and recognize disproportionation.

Slide 1 of 15 · A model for shared electrons

Charge equals the number of protons **minus** the number of electrons. Na^+ has 11 protons and 10 electrons, so its charge is **+1**.

For a one-atom ion, the oxidation number equals the **ion charge**. It tells how many electrons the atom gained or lost compared with the neutral atom.

In HCl, chlorine pulls the shared pair harder. This is greater **electronegativity**. We give both bonding electrons to chlorine, so H is +1 and Cl is **−1**.

If an atom's oxidation number goes **up**, it is oxidized. If it goes **down**, it is reduced. This works even when both numbers are positive, such as +5 to +2.

Example. H_2 and HCl are molecules that share electrons. Chlorine is more electronegative than hydrogen. Find the oxidation number of each atom in H_2 and in HCl.

1. In H_2 the two atoms are identical, so they share the pair equally. Each H gets one electron, so each H is 0.

2. In HCl both bonding electrons go to Cl. H has one fewer electron than a neutral H atom, so H is +1.

3. Cl has one extra electron, so Cl is −1.

4. Check: (+1) + (−1) = 0, which fits a neutral molecule.

Try it. A magnesium ion Mg^{2+} has 12 protons and 10 electrons. What is its charge, and what is its oxidation number?

1. Charge = protons − electrons = 12 − 10 = +2, which is written 2+.

2. For a one-atom ion the oxidation number equals the charge, so Mg is +2.

Slide 2 of 15 · Read the formula before assigning values

A small lower number, called a **subscript**, counts atoms in one particle. No subscript means the count is **one**.

A big number in front, called a **coefficient**, counts copies of the whole formula. In $2SO_4^{2-}$ there are **two** sulfate ions.

Brackets keep a group together. In $Ca(ClO_4)_2$ the 2 outside multiplies every atom inside, so there are **2** Cl and **8** O but only 1 Ca.

An ion charge is written **number then sign**, as in 2−. An oxidation number is written **sign then number**, as in +6.

Example. Compare $2SO_4^{2-}$ and $Ca(ClO_4)_2$. For each, give the number of atoms of every element and say what charge the formula carries.

1. $2SO_4^{2-}$: the coefficient 2 means two sulfate ions. That is 2 S and 2 × 4 = 8 O.

2. Each sulfate ion has charge 2−, so the two ions carry 2(2−) = 4− in total.

3. $Ca(ClO_4)_2$: 1 Ca, 2 Cl and 2 × 4 = 8 O.

4. There is no raised charge, so $Ca(ClO_4)_2$ is neutral (total charge 0).

Try it. Count the atoms of each element in $Al_2(SO_4)_3$ and give its total charge.

1. The 2 after Al counts 2 Al atoms.

2. The 3 outside the brackets multiplies S and O: 3 S and 3 × 4 = 12 O.

3. There is no raised charge, so the formula is neutral (total charge 0).

Slide 3 of 15 · Rules that fix the known values

An element by itself, such as Cl_2 or Na, has an oxidation number of **0**. Check this before any compound rule.

A monatomic ion has an oxidation number equal to its **charge**. For example, Al^{3+} is **+3**.

In compounds, Group 1 metals are usually **+1** and Group 2 metals are **+2**. Oxygen is usually **−2**, hydrogen is usually +1 and fluorine is always **−1**.

The oxidation numbers of all atoms must add up to the **charge of the whole species**.

Example. Given Cl_2, Al^{3+} and MgO. Use the rules to find the oxidation number of every atom without doing any algebra.

1. Cl_2 is an element by itself, so Cl is 0.

2. Al^{3+} is a monatomic ion, so Al is +3.

3. In MgO, Mg is Group 2, so Mg is +2, and O is −2.

4. Check MgO: (+2) + (−2) = 0, which fits a neutral compound.

Try it. Given Na, Cl^- and CaF_2. Assign an oxidation number to every atom using the rules in order.

1. Na is an element by itself, so Na is 0.

2. Cl^- is a monatomic ion, so Cl is −1.

3. In CaF_2, Ca is Group 2, so Ca is +2, and F is always −1.

4. Check CaF_2: (+2) + 2(−1) = 0.

Slide 4 of 15 · Oxidation-number rules · reference

Check the rules in order: an element by itself is **0**, a one-atom ion equals its **charge**, then the compound rules.

Hydrogen is usually **+1**, but it is −1 in a metal **hydride**.

Oxygen is usually −2, but it is **−1** in a peroxide. It can be positive when bonded to **fluorine**.

The sum of all atom values is **0** for a neutral compound. For a polyatomic ion the sum is the **ion charge**.

Example. Use the rules in order to find the oxidation number of every atom in K_2O. Name the rule you used for each value.

1. K is Group 1 in a compound, so K is +1.

2. O is usually −2, so O is −2.

3. Check: 2(+1) + (−2) = 0, which is the total for a neutral compound.

Try it. Use the rules to find the oxidation number of every atom in CaH_2. Name the rule you used for hydrogen.

1. Ca is Group 2 in a compound, so Ca is +2.

2. CaH_2 is a metal hydride, so H is −1.

3. Check: (+2) + 2(−1) = 0.

Slide 5 of 15 · Start with neutral compounds

In KCl, potassium is +1 and chlorine is **−1**, so the sum is **0**. Chlorine is 0 only when it is an element by itself, as in Cl_2.

In CO, oxygen is −2. Let x be carbon. Then x + (−2) = 0, so carbon is **+2**.

In NH_3, three hydrogens give 3(+1) = **+3**. The total must be 0, so nitrogen is **−3**.

A subscript is an atom **count**. It multiplies the value. It does not change the oxidation number of one atom.

Example. Find the oxidation number of every atom in KCl, CO and NH_3. Start with a known rule, then use the total charge of the formula.

1. KCl: K is +1 (Group 1) and Cl is −1. Check: (+1) + (−1) = 0.

2. CO: O is −2. x + (−2) = 0, so C is +2.

3. NH_3: each H is +1, so three H give +3. x + 3(+1) = 0, so N is −3.

4. Check: −3 + 1 + 1 + 1 = 0.

Try it. Find the oxidation number of nitrogen in NO and of sulfur in H_2S.

1. NO: O is −2. x + (−2) = 0, so N is +2.

2. H_2S: each H is +1, so two H give +2. x + 2(+1) = 0, so S is −2.

3. Check H_2S: 2(+1) + (−2) = 0.

Slide 6 of 15 · A complete example: sulfur in sulfate

The 2− on SO_4^{2-} belongs to the **whole ion**. It is the target total, **−2**. It is not the value for sulfur or for each oxygen.

Four oxygens at −2 give 4(−2) = **−8**. Let x be sulfur. Then x + 4(−2) = −2, so x = **+6**.

Check by putting the answer back: +6 + 4(−2) = **−2**. The +6 belongs to sulfur alone. The 2− is the charge of the whole **ion**.

For NH_4^+ the target is **+1**, not zero. So x + 4(+1) = +1 and nitrogen is **−3**.

Example. Given the sulfate ion SO_4^{2-}. Find the oxidation number of sulfur and show the counting.

1. Oxygen is −2, so four oxygens give 4(−2) = −8.

2. The target is the ion charge, −2. Let x be sulfur: x + 4(−2) = −2.

3. x − 8 = −2, so x = +6.

4. Check: +6 + 4(−2) = +6 − 8 = −2.

Try it. Given the ammonium ion NH_4^+. Find the oxidation number of nitrogen. The ion has one N and four H.

1. Hydrogen is +1, so four hydrogens give 4(+1) = +4.

2. The target is the ion charge, +1. Let x be nitrogen: x + 4(+1) = +1.

3. x + 4 = +1, so x = −3.

4. Check: −3 + 4 = +1.

Slide 7 of 15 · Method: how to find an oxidation number

Step 1. Write the formula and its overall **charge**. A neutral compound is **0**.

Step 2. Write the **known** values from the rules, such as O usually −2 and H usually +1.

Step 3. Let **x** stand for the unknown atom.

Step 4. Set the **sum** of all the oxidation numbers equal to the overall charge.

Step 5. **Solve** for x and include the sign.

Check: add the numbers again. The total must equal the **overall charge**.

Example. Use the method on SO_4^{2-}. Find the oxidation number of sulfur.

1. Step 1: the overall charge is −2.

2. Step 2: oxygen is usually −2.

3. Steps 3 and 4: x + 4(−2) = −2.

4. Step 5: x = +6. Check: +6 − 8 = −2.

Try it. Use the method on NO_3^-. Find the oxidation number of nitrogen.

1. Overall charge is −1 and oxygen is −2.

2. x + 3(−2) = −1, so x = +5.

3. Check: +5 − 6 = −1.

Slide 8 of 15 · When the unknown element appears twice

$H_2P_2O_7^{2-}$ has 2 H, 2 P and 7 O, and its charge is **2−**. The hydrogens give **+2** and the oxygens give 7(−2) = −14.

The known atoms give +2 − 14 = −12. The ion must total −2, so the two phosphorus atoms must give **+10** together.

Write 2x for two phosphorus atoms: 2(+1) + 2x + 7(−2) = −2. Then 2x = 10 and x = **+5**.

The question asks for one atom, so divide by the **count** at the end. The answer is an **average** value for phosphorus.

Example. Given the ion $H_2P_2O_7^{2-}$. Find the average oxidation number of phosphorus and check the total for the whole ion.

1. H is +1, so 2 H give +2. O is −2, so 7 O give −14.

2. Let x be one phosphorus atom: 2(+1) + 2x + 7(−2) = −2.

3. 2x − 12 = −2, so 2x = +10 and x = +5.

4. Check: +2 + 2(+5) − 14 = −2.

Try it. Given the thiosulfate ion $S_2O_3^{2-}$. Find the average oxidation number of sulfur.

1. O is −2, so 3 O give −6.

2. Let x be one sulfur atom: 2x + 3(−2) = −2.

3. 2x − 6 = −2, so 2x = +4 and x = +2.

4. Check: 2(+2) − 6 = −2.

Slide 9 of 15 · Brackets change the counts, not the method

In $Ca(ClO_4)_2$ the 2 outside the brackets multiplies every atom inside. There is 1 Ca, **2** Cl and **8** O.

Calcium is +2 and oxygen is −2. The formula has no charge, so the total is **0**.

The sum is $+2 + 2x + 8(-2) = 0$. Then $2x = 14$, so chlorine is **+7**. The +14 is the total for two atoms, not the answer.

Check one ClO_4 group: $+7 + 4(-2) =$ **−1**. Two such groups give −2, which balances the +2 of calcium.

Example. Given the compound $Ca(ClO_4)_2$. Find the oxidation number of chlorine. Show how the outside 2 affects every atom in the brackets.

1. Count atoms: 1 Ca, 2 Cl and $2 \times 4 = 8$ O.

2. Ca is +2 and O is −2. The formula is neutral, so the total is 0.

3. $+2 + 2x + 8(-2) = 0$, so $2x - 14 = 0$ and $x = +7$.

4. Check: $+2 + 2(+7) + 8(-2) = +2 + 14 - 16 = 0$.

5. One ClO_4 group: $+7 + 4(-2) = -1$.

Try it. Given the compound $Mg(NO_3)_2$. Find the oxidation number of nitrogen. Count the atoms first.

1. Count atoms: 1 Mg, 2 N and $2 \times 3 = 6$ O.

2. Mg is +2 and O is −2. The formula is neutral, so the total is 0.

3. $+2 + 2x + 6(-2) = 0$, so $2x - 10 = 0$ and $x = +5$.

4. Check one NO_3 group: $+5 + 3(-2) = -1$.

Slide 10 of 15 · Your turn: nitrate

Start by counting atoms. NO_3^- has **3** oxygens, and each oxygen is **−2**.

The target total is the ion charge, **−1**, not zero.

Write one equation with x for nitrogen: x + 3(−2) = −1, so x = **+5**.

Check by putting the answer back: +5 − 6 = **−1**. If you wrote x − 6 = 0, the **target** was wrong. If you wrote x − 2 = −1, two **oxygens** were missing.

Example. Given the nitrate ion NO_3^-. Find the oxidation number of nitrogen. Write your sum with signs.

1. Three oxygens at −2 give 3(−2) = −6.

2. The target is the ion charge, −1. Let x be nitrogen: x + 3(−2) = −1.

3. x − 6 = −1, so x = +5.

4. Check: +5 − 6 = −1.

Try it. Given the carbonate ion CO_3^{2-}. Find the oxidation number of carbon. Write your sum with signs.

1. Three oxygens at −2 give 3(−2) = −6.

2. The target is the ion charge, −2. Let x be carbon: x + 3(−2) = −2.

3. x − 6 = −2, so x = +4.

4. Check: +4 − 6 = −2.

Slide 11 of 15 · Exceptions and what an average means

In a peroxide such as H_2O_2, each oxygen is **−1**. With H at +1, 2(+1) + 2x = 0.

In a metal hydride such as NaH, sodium is +1, so hydrogen is **−1**.

Fluorine is always **−1**, so in OF_2 oxygen is **+2**. Assign fluorine first, then oxygen.

When the sum gives a fraction, as in C_3H_8 where 3x + 8 = 0, x is **−8/3**. This is an **average**, so do not round it away.

Example. Compare H_2O_2, NaH and OF_2. Find the oxidation number of oxygen in H_2O_2 and OF_2 and of hydrogen in NaH. Say which usual rule changes each time.

1. H_2O_2 (peroxide): H is +1. 2(+1) + 2x = 0, so O is −1. Oxygen is not −2.

2. NaH (metal hydride): Na is +1. +1 + x = 0, so H is −1. Hydrogen is not +1.

3. OF_2: F is −1. x + 2(−1) = 0, so O is +2.

Try it. BaO_2 is a peroxide. Find the oxidation number of oxygen in BaO_2 and the average oxidation number of carbon in C_4H_{10}.

1. BaO_2: Ba is +2 (Group 2). +2 + 2x = 0, so each O is −1.

2. C_4H_{10}: H is +1, so ten H give +10. 4x + 10(+1) = 0, so 4x = −10.

3. x = −10/4 = −5/2, which is −2.5. This is the average for one carbon atom.

Slide 12 of 15 · A covalent reaction is still redox

In CH_4, four H at +1 total +4, so carbon is **−4**. In CO_2, two O at −2 total −4, so carbon is **+4**.

Carbon goes from −4 to +4, an increase of **8**, so carbon is **oxidized**. Each oxygen atom goes from 0 to −2, so oxygen is reduced.

The balanced equation has $2O_2$, which is 4 O atoms. Four atoms each gaining 2 give **8** electrons gained, which matches the carbon loss.

Name the agent as a whole reactant. CH_4 is the **reducing** agent and O_2 is the **oxidizing** agent.

Example. Given $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$. Identify the oxidized element, the reduced element and both reactant agents.

1. CH_4: $4(+1) + x = 0$, so C is −4. CO_2: $x + 2(-2) = 0$, so C is +4.

2. C rises from −4 to +4, so carbon is oxidized.

3. O_2 is an element by itself, so O is 0. In the products O is −2, so oxygen is reduced.

4. Electrons: C loses 8. Four O atoms each gain 2, which is 8. They match.

5. CH_4 contains the oxidized carbon, so CH_4 is the reducing agent. O_2 is the oxidizing agent.

Try it. Given $S(s) + O_2(g) \rightarrow SO_2(g)$. Find the oxidation numbers, the electrons lost and gained, and both agents.

1. S and O_2 are elements by themselves, so S is 0 and O is 0.

2. In SO_2, O is −2, so $x + 2(-2) = 0$ and S is +4.

3. S rises 0 to +4 and loses 4 electrons, so S is oxidized and is the reducing agent.

4. Two O atoms each fall 0 to −2 and gain 4 electrons, so O_2 is reduced and is the oxidizing agent.

Slide 13 of 15 · Method: how to find what was oxidized and reduced

Step 1. **Number** every atom on both sides of the equation.

Step 2. **Compare** each element before and after.

Step 3. If the number goes **up**, the particle is oxidized.

Step 4. If the number goes **down**, the particle is reduced.

Step 5. Count the **electrons** for each change and check they match.

Check: the electrons **lost** must equal the electrons **gained**.

Example. Use the method on $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$. Name what is oxidized and what is reduced.

1. Step 1: C is -4 in CH_4 and $+4$ in CO_2. O is 0 in O_2 and -2 in CO_2. H is $+1$ on both sides.

2. Steps 3 and 4: carbon goes up, so CH_4 is oxidized. Oxygen goes down, so O_2 is reduced.

3. Step 5: carbon loses 8 e^-. Four O atoms gain $4 \times 2 = 8$ e^-.

Try it. Use the method on $Zn + Cu^{2+} \rightarrow Zn^{2+} + Cu$. Name what is oxidized and what is reduced.

1. Zn goes from 0 to $+2$, so Zn is oxidized.

2. Cu goes from $+2$ to 0, so Cu^{2+} is reduced.

3. Two electrons are lost and two are gained.

Slide 14 of 15 · Your turn: hydrogen and oxygen

In H_2 and O_2 each atom is **0**, because each is an element by itself.

In water, hydrogen is **+1** and oxygen is **−2**.

Coefficients count whole molecules. $2H_2$ has 4 H atoms, each up by 1, so **4** electrons are lost. O_2 has 2 O atoms, each down by 2, so **4** electrons are gained.

Hydrogen is oxidized, so H_2 is the **reducing** agent. Oxygen is reduced, so O_2 is the oxidizing agent.

Example. Given $2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$. Find the oxidation numbers of H and O before and after, name both agents and count the total electrons transferred.

1. H_2 and O_2 are elements by themselves, so H is 0 and O is 0.

2. In H_2O, H is +1 and O is −2 (check: $2(+1) + (-2) = 0$).

3. H goes 0 to +1, so H_2 is oxidized and is the reducing agent. O goes 0 to −2, so O_2 is reduced and is the oxidizing agent.

4. Four H atoms each lose 1, which is 4. Two O atoms each gain 2, which is 4. So 4 electrons are transferred.

Try it. Given $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$. Assign N and H before and after, name both agents and count the total electrons transferred.

1. N_2 and H_2 are elements by themselves, so N is 0 and H is 0.

2. In NH_3, H is +1, so $x + 3(+1) = 0$ and N is −3.

3. Two N atoms each go 0 to −3 and gain 6 electrons, so N_2 is reduced and is the oxidizing agent.

4. Six H atoms each go 0 to +1 and lose 6 electrons, so H_2 is oxidized and is the reducing agent.

5. Total electrons transferred: 6.

Slide 15 of 15 · One reactant can do both jobs

In $2H_2O_2 \rightarrow 2H_2O + O_2$, oxygen is -1 in H_2O_2, **-2** in H_2O and **0** in O_2.

Oxygen going to -2 is **reduced**. Oxygen going to 0 is oxidized. Some atoms go each way, so follow both destinations.

One reactant that is both oxidized and reduced shows **disproportionation**. Here H_2O_2 is both the oxidizing agent and the **reducing** agent.

If no oxidation number changes, as in $CaCO_3 \rightarrow CaO + CO_2$, the reaction is **not** redox.

Example. Given $2H_2O_2(aq) \rightarrow 2H_2O(l) + O_2(g)$. Is this redox? Compare the oxidation number of oxygen in all three species and name the agents.

1. H_2O_2: H is +1, so $2(+1) + 2x = 0$ and O is -1. In H_2O, O is -2. In O_2, O is 0.

2. O goes -1 to -2 (reduced) and -1 to 0 (oxidized). Oxidation numbers change, so it is redox.

3. H_2O_2 is both oxidized and reduced, so it is both the oxidizing agent and the reducing agent. This is disproportionation.

Try it. Given $CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$. Is this redox? Find the oxidation numbers of Ca, C and O in every formula and justify your answer.

1. $CaCO_3$: Ca is +2, O is -2 and $2 + x + 3(-2) = 0$, so C is +4.

2. CaO: Ca +2, O -2. CO_2: C is +4, O is -2.

3. No oxidation number changes, so no electrons transfer. It is not redox.

B2 · Worked practice

Oxidation numbers and redox changes | Questions 1-2 of 7

1. Find the oxidation number of Mn in MnO_4^-. Write the sum of all the oxidation numbers, with signs. Set it equal to the charge of the ion.

Use Method: how to find an oxidation number. x + 4(−2) = −1, so x = +7. Check: +7 − 8 = −1.

2. Find the oxidation number of Cr in $Cr_2O_7^{2-}$. Explain why x is multiplied by 2 in your sum.

Use Method: how to find an oxidation number. 2x + 7(−2) = −2, so 2x = +12 and x = +6. There are two Cr atoms, so x is multiplied by 2.

B2 · Worked practice

Oxidation numbers and redox changes | Questions 3-4 of 7

3. Find the oxidation number of oxygen in O_2, H_2O_2 and OF_2. Explain why the three values are different.

In O_2 it is 0, because O_2 is an element by itself. In H_2O_2 it is -1, because H_2O_2 is a peroxide. In OF_2 it is $+2$, because F is always -1. The usual value of -2 does not fit these three.

4. Given $2CO(g) + O_2(g) \rightarrow 2CO_2(g)$. Name both agents. How many electrons move in total?

Use Method: how to find what was oxidized and reduced. C is $+2$ in CO and $+4$ in CO_2. So CO is oxidized, and CO is the reducing agent. O in O_2 goes from 0 to -2. So O_2 is reduced, and O_2 is the oxidizing agent. Two C atoms each go up by 2, so 4 electrons move.

B2 · Worked practice

Oxidation numbers and redox changes | Questions 5-6 of 7

5. Given $HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)$. Use oxidation numbers to decide if this is a redox reaction.

No. H stays +1, Cl stays −1, Na stays +1 and O stays −2. No oxidation number changes, so this reaction of an acid with a base is not redox.

6. Given $2H_2O_2 \rightarrow 2H_2O + O_2$. Which reactant is both agents? Explain.

H_2O_2 is both agents. Its oxygen starts at −1. Some of it is reduced to −2 in water, and some is oxidized to 0 in O_2. This is disproportionation.

B2 · Worked practice

Oxidation numbers and redox changes | Questions 7-7 of 7

7. Challenge: Given the ions $S_2O_3^{2-}$ and $S_4O_6^{2-}$. Find the average oxidation number of sulfur in each ion. Then say whether sulfur is oxidized or reduced when $S_2O_3^{2-}$ changes to $S_4O_6^{2-}$.

In $S_2O_3^{2-}$: $2x + 3(-2) = -2$, so $2x = +4$ and $x = +2$. In $S_4O_6^{2-}$: $4x + 6(-2) = -2$, so $4x = +10$ and $x = +2.5$. The average is not a whole number because the sulfur atoms are not all alike. The number rises from +2 to +2.5, so sulfur is oxidized.