B2 · Oxidation numbers and redox changes

Completed notes | Notes 1/3

Find the oxidation number of each atom with the rules and the sum rule. Know the exceptions. Use the numbers to find what is oxidized and reduced.

1. Rules and formula counting

An assigned value, not always a real ion charge

An oxidation number is an **electron-counting assignment**. It is the charge an atom would have if we gave all the bonding electrons to the more electronegative atom, the one that pulls on bonding electrons harder. Identical atoms share the electrons equally.

In H_2 the two atoms are the same, so each H is 0. In HCl, chlorine pulls harder, so Cl is -1 and H is +1. This is a counting method, not a measurement. HCl gas does not contain separate H^+ and Cl^- ions.

An element by itself has oxidation number **0**, including O_2 and F_2. A one-atom ion, called a monatomic ion, has oxidation number equal to **its ion charge**. For Na^+: 11 protons - 10 electrons = +1.

On the diploma: Cl_2 is 0, not -1, because it is an element by itself. Giving it the usual compound value is a common wrong option.

Watch: The oxidation number rules (Tyler DeWitt, “How to Calculate Oxidation Numbers Introduction”, 0:26)

Watch: An element by itself is 0 (Tyler DeWitt, “How to Calculate Oxidation Numbers Introduction”, 0:58)

Watch: A one-atom ion equals its charge (Tyler DeWitt, “How to Calculate Oxidation Numbers Introduction”, 1:47)

Keep the rule set together

Check the free-element and one-atom-ion cases first. Then use the compound rules. In compounds, Group 1 is **+1**, Group 2 is **+2**, and F is **-1**. Other halogens are usually **-1**, but Cl, Br and I can be positive when combined with O or F.

H is usually **+1**, but it is **-1** in metal hydrides such as NaH. O is usually **-2**, but it is **-1** in peroxides and can be positive with fluorine. These are usual values, so check each formula for an exception.

The oxidation numbers of all atoms must add up to the charge of the whole formula. A neutral compound has sum **0**. A polyatomic ion, a charged group of bonded atoms, has sum equal to **the whole ion’s signed charge**. A subscript multiplies an atom’s value, but a coefficient in front does not change it.

On the diploma: write oxidation numbers with the sign first, such as +6. Numerical-response answers record only the digits asked for, so read how to enter the sign.

Watch: Group 1A, group 2A and the halogens (Tyler DeWitt, “How to Calculate Oxidation Numbers Introduction”, 2:49)

Watch: Hydrogen is −1 with a metal (Tyler DeWitt, “How to Calculate Oxidation Numbers Introduction”, 9:45)

Watch: Oxygen is −1 in a peroxide (Tyler DeWitt, “How to Calculate Oxidation Number Practice Problems”, 9:49)

Watch: Halogens are positive with oxygen (Tyler DeWitt, “How to Calculate Oxidation Numbers Introduction”, 10:36)

Watch: Fluorine is always −1 (Tyler DeWitt, “How to Calculate Oxidation Numbers Introduction”, 11:50)

Build the sum

CO: oxygen is -2 and CO is neutral, so x + (-2) = 0. That makes carbon **+2**. We find carbon’s value from its own formula, not from another compound.

NH_3: each H is +1, and three H contribute 3 × (+1) = **+3**. The total must be 0, so nitrogen is **-3**. The subscript 3 counts atoms. It does not make each H +3.

For the sulfate diagram, label three things: which number is an atom count, which is the whole-ion charge, and which is the per-oxygen value. Then complete the sulfur calculation.

Watch: A neutral compound adds to 0 (Tyler DeWitt, “How to Calculate Oxidation Numbers Introduction”, 3:31)

Watch: Find the unknown: carbon in CO (Tyler DeWitt, “How to Calculate Oxidation Numbers Introduction”, 4:53)

Watch: An ion adds to its charge (Tyler DeWitt, “How to Calculate Oxidation Numbers Introduction”, 7:45)

Watch: Diploma example: sulfur in H_2SO_3 (Alberta Chemistry Teacher, “Electrochemistry Diploma Prep - Chemistry 30 - Part 1 of 4”, 12:57)

Given SO_4^{2-}; let sulfur = x.

O atom count: 4

Whole-ion charge: -2

Each O: -2

All O total: -8

Sum and sulfur: x + 4(-2) = -2 → x = +6

B2 · Oxidation numbers and redox changes

Completed notes | Notes 2/3

2. Read the formula, then solve step by step

Read the whole formula first

Before you assign anything, count what is in the formula. In SO_4^{2-} there is **one** S and **four** O. The whole ion has a charge of **-2**. In $2SO_4^{2-}$ you have two copies, but each ion still has a charge of **-2**.

Brackets keep a group together. In $Ca(ClO_4)_2$ the 2 outside multiplies **both Cl and O inside the parentheses**. So the full count is **one Ca, two Cl and eight O**. No charge is written, so the atoms must add up to **0**.

On the diploma: count the atoms inside brackets first. Most wrong answers come from forgetting that the outside number multiplies everything inside.

Watch: Multiply by the number of atoms (Tyler DeWitt, “How to Calculate Oxidation Number Practice Problems”, 2:28)

Two signed values, written two ways

An ion charge is written with the number first, then the sign: Ca^{2+} is **2+** and SO_4^{2-} is **2−**. An oxidation number is written with the sign first, then the number: calcium is **+2** and sulfur in sulfate is **+6**.

The charge belongs to the **whole ion**. The oxidation number is a value we assign to **one atom**. Try NH_4^+. Hydrogen is +1, so x + 4(+1) = +1. That makes N equal to **-3**. Check: -3 + 4 = **+1**, which matches the ion charge.

Watch: Write the sign first for oxidation numbers (Tyler DeWitt, “How to Calculate Oxidation Numbers Introduction”, 2:18)

Watch: Charge and oxidation number: number and sign swap places (Alberta Chemistry Teacher, “Electrochemistry Diploma Prep - Chemistry 30 - Part 1 of 4”, 22:56)

When the unknown element appears twice

Take $H_2P_2O_7^{2-}$. Hydrogen is +1 and oxygen is -2. The known atoms add to 2(+1) + 7(-2) = **-12**. The ion must total -2, so the two phosphorus atoms together must supply **+10**.

Write it as an equation, with x as the value for one P: 2(+1) + 2x + 7(-2) = -2. Solve it: 2x = **+10**, so x = **+5**. Check: +2 + 2(+5) - 14 = **-2**.

On the diploma: if the unknown element appears twice, divide your total by two at the end. Then check the sum against the ion charge.

Watch: Two carbons in CaC_2 (Tyler DeWitt, “How to Calculate Oxidation Numbers Introduction”, 6:58)

Watch: Two phosphorus atoms (Tyler DeWitt, “How to Calculate Oxidation Number Practice Problems”, 6:11)

A full worked example, then your turn

Let’s do $Ca(ClO_4)_2$ with Ca as +2 and O as -2. The equation is +2 + 2x + 8(-2) = 0. Solve: 2x = **14**, so each Cl is **+7**. Check: +2 + 14 - 16 = **0**.

Look at one ClO_4^- group on its own: +7 + 4(-2) = **-1**. Two of those groups balance one Ca^{2+}, so the Cl value stays the same in each group.

Your turn with $Al_2(SO_4)_3$. Take Al as +3 and O as -2. Count first: two Al, **three S** and **twelve O**. Then 2(+3) + 3x + 12(-2) = 0, so each S is **+6**. Check: +6 + 18 - 24 = **0**.

Watch: A formula with brackets: $Ca(ClO_4)_2$ (Tyler DeWitt, “How to Calculate Oxidation Number Practice Problems”, 12:55)

B2 · Oxidation numbers and redox changes

Completed notes | Notes 3/3

3. Exceptions, averages and reaction changes

Model and check

In SO_4^{2-} the target is the charge printed on the ion, -2. So x + 4(-2) = -2, and sulfur is **+6**. Check: +6 - 8 = **-2**. The +6 belongs to sulfur alone. Sulfate does not contain a free S^{6+} ion.

H_2O_2: with H at +1, 2(+1) + 2x = 0 gives O = **-1**. NaH: +1 + x = 0 gives H = **-1**. OF_2: F is -1, so x + 2(-1) = 0 gives O = **+2**. Give fluorine its value first. The sum rule still works when a usual value changes.

In C_3H_8 the eight H at +1 give +8, so 3x + 8(+1) = 0. Average carbon is **-8/3**. This is an average, not a real fractional charge, so do not round it away.

On the diploma: wrong options are often built from a wrong target. Use the printed ion charge, never zero, for an ion such as sulfate.

Watch: An ion: CO_3^{2-} (Tyler DeWitt, “How to Calculate Oxidation Number Practice Problems”, 4:37)

Compare reactants with products

An increase in oxidation number means **oxidation**. A decrease means **reduction**. This works even when both numbers are positive, such as +5 to +2. In $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$, carbon changes **-4 to +4** and oxygen changes **0 to -2**.

Check that the counting balances. Carbon rises by 8. Four oxygen atoms each fall by 2, and 4 × 2 = 8. So the electrons lost equal the electrons gained.

The reducing agent is the whole starting substance **CH_4**. The oxidizing agent is **O_2**. Find the changing element, then go back to its reactant formula. Naming the element or a product is a common wrong choice.

In $2H_2O_2 \rightarrow 2H_2O + O_2$, oxygen starts at -1 and ends at **-2 and 0**. The same reactant is both oxidized and reduced. This is **disproportionation**. Some peroxide oxygen goes down to water and some goes up to O_2.

On the diploma: to call a reaction "not redox", say that no oxidation number changes.

Watch: Number goes up is oxidation, down is reduction (Tyler DeWitt, “Oxidizing Agents and Reducing Agents”, 6:13)

Watch: Diploma example: what is oxidized in aluminium and chlorine (Alberta Chemistry Teacher, “Electrochemistry Diploma Prep - Chemistry 30 - Part 1 of 4”, 2:47)

Watch: Diploma example: oxidation numbers in a combustion reaction (Alberta Chemistry Teacher, “Electrochemistry Diploma Prep - Chemistry 30 - Part 1 of 4”, 8:32)

B2 · Worked practice

Oxidation numbers and redox changes | Questions 1-2 of 12

State symbols: (s) solid, (aq) dissolved in water, (l) liquid, (g) gas. Use the supplied reaction and conditions. B4 values name the Alberta Chemistry 30 Data Booklet table where used.

B2-P01 · Easy · Find the oxidation numbers of Mg(s), O_2(g), Fe^{3+}(aq) and Cl^-(aq). Say which rule you used for each one.

1. Mg and O_2 are elements by themselves, so each is 0. The subscript in O_2 does not make oxygen -2.

2. Fe^{3+} is +3 and Cl^- is -1. For an ion made of one atom, the oxidation number is the charge.

B2-P02 · Easy · Find the oxidation numbers of K and Cl in KCl. Show that they add up to the charge of the formula.

1. K is +1 because it is a Group 1 metal. Cl is -1 here, the usual value for a halogen.

2. One of each gives +1 + (-1) = 0. KCl is neutral, so the total must be 0.

B2 · Worked practice

Oxidation numbers and redox changes | Questions 3-4 of 12

B2-P03 · Easy · Find the oxidation number of carbon in CO_2. Show how the subscript on oxygen is used in the sum.

1. Each oxygen is -2, so two oxygens give 2(-2) = -4. Let carbon be x.

2. x - 4 = 0, so x = +4. The total is 0 because CO_2 is neutral. Each O is still -2.

B2-P04 · Easy · Find the oxidation number of nitrogen in NH_3. Give the value for one H, and the total for all the H atoms.

1. Each H is +1. Three hydrogens give 3(+1) = +3.

2. x + 3 = 0, so N = -3. Do not give +3 to each H. The 3 counts the atoms.

B2 · Worked practice

Oxidation numbers and redox changes | Questions 5-6 of 12

B2-P05 · Medium · Find the oxidation number of sulfur in SO_3^{2-}. Write the total the numbers must add to before you solve. Then check your answer.

1. The charge of the whole ion is -2, so the numbers must add to -2. Each O is -2, so x + 3(-2) = -2.

2. x - 6 = -2, so S = +4. Check by putting it back: +4 - 6 = -2.

B2-P06 · Medium · Find the oxidation number of nitrogen in NO_3^- and then in NO_2^-. If nitrate changes to nitrite, is nitrogen oxidized or reduced?

1. Nitrate: x + 3(-2) = -1, so x = +5. Nitrite: x + 2(-2) = -1, so x = +3.

2. The number goes down from +5 to +3, so nitrogen is reduced. The charge of the ion stays -1, but the oxidation number of nitrogen changed.

B2 · Worked practice

Oxidation numbers and redox changes | Questions 7-8 of 12

B2-P07 · Medium · Find the oxidation number of oxygen in H_2O_2 and in OF_2. Say why the usual value for oxygen does not work in each one.

1. H_2O_2 is a peroxide. With H at +1, 2 + 2x = 0, so O = -1.

2. In OF_2, F is always -1. So x - 2 = 0, and O = +2. The usual -2 for oxygen does not work in either one.

B2-P08 · Medium · Find the oxidation number of hydrogen in CaH_2. Then find the oxidation number of bromine in BrO_3^-. Show both sums.

1. Ca is a Group 2 metal, so +2 + 2x = 0 and each H = -1. This is a metal hydride, one of the exceptions.

2. In BrO_3^-, oxygen is -2, so x - 6 = -1 and Br = +5. A halogen can be positive when it is joined to oxygen.

B2 · Worked practice

Oxidation numbers and redox changes | Questions 9-10 of 12

B2-P09 · Challenging · For $C_2H_6 + 7/2\ O_2 \rightarrow 2CO_2 + 3H_2O$, find the average oxidation number of carbon on each side. What is oxidized and what is reduced? Name both agents. How many electrons does each C_2H_6 lose?

1. In C_2H_6, $2x + 6 = 0$, so the average for C is -3. In CO_2, C is +4. Each carbon goes up by 7, so the two carbon atoms lose 14 electrons in total.

2. O goes down from 0 to -2, so it is reduced. C_2H_6 is the reducing agent and O_2 is the oxidizing agent. Name the whole reactants, not a single carbon or oxygen atom.

3. Seven oxygen atoms each gain two electrons, so 14 are gained. That matches the 14 lost. The 7/2 in front of O_2 is only a ratio. Doubling every coefficient gives whole numbers.

B2-P10 · Challenging · A student solves NO_3^- with $x + 3(-2) = 0$ and gets +6. Find the mistake and fix the sum. Explain why checking the arithmetic alone would not find the mistake.

1. The student set the total to 0, but the nitrate ion has a charge of -1. The arithmetic is right, but the sum is the wrong one.

2. Use $x - 6 = -1$, so N = +5. Always check the total against the charge of the ion. Not every formula is neutral.

B2 · Worked practice

Oxidation numbers and redox changes | Questions 11-12 of 12

B2-P11 · Challenging · For $2H_2O_2 \rightarrow 2H_2O + O_2$, find the oxidation number of oxygen in all three substances. Explain why H_2O_2 is both agents. Explain why having two products does not prove disproportionation.

1. O is -1 in H_2O_2, -2 in H_2O and 0 in O_2. Some oxygen goes down from -1 and some goes up from -1.

2. H_2O_2 is both the oxidizing agent and the reducing agent, because the same reactant is reduced and oxidized. Two products alone would not show those two changes.

B2-P12 · Challenging · Find the average oxidation number of carbon in C_2H_6O. A student says the answer is the real charge on each carbon atom. Explain what the calculation shows and what it does not show.

1. Use H = +1 and O = -2: 2x + 6(+1) - 2 = 0. So 2x = -4, and the average for C is -2.

2. This is a count for the whole formula, averaged over two carbon atoms. It does not show that the two carbons are the same. It does not show real C^{2-} ions in the molecule. Looking at each carbon by itself is not part of this lesson.