Khan Academy

# High School Chemistry

## Unit 4: Chemical reactions

### Overview

In this unit, students will explore different types of chemical reactions and learn how to represent, analyze, and predict their outcomes.

**Lesson 1:** Students will write and interpret chemical equations representing changes from reactants to products in chemical reactions.

**Lesson 2:** Students will connect the law of conservation of mass to the rearrangement of atoms in a chemical reaction and use coefficients to balance chemical equations.

**Lesson 3:** Hands-on science activity (see below)

**Lesson 4:** Students will **develop and use models** for energy changes in chemical reaction systems and describe reactions as exothermic or endothermic.

**Lesson 5:** Students will analyze reaction patterns to classify reactions by type (synthesis, decomposition, combustion, or single replacement), to predict products, and to write balanced chemical equations.

### Hands-on science activity

***How can a big log turn into a tiny pile of ash when it burns?***

Students will **carry out an investigation** and **analyze and interpret data** to determine what happens to the amount of matter during a physical or chemical change. They will use their experimental data as evidence to support the claim that mass is conserved. Click here for links to the activity.

### Standards

Performance expectations: **HS-PS1-2 | HS-PS1-4 | HS-PS1-7**

Disciplinary core ideas: **HS-PS1.A.4 | HS-PS1.B.1 | HS-PS1.B.3**

Science and engineering practices:

- Developing and using models
- Planning and carrying out investigations
- Engaging in argument from evidence
- Analyzing and interpreting data

Crosscutting concepts:

- Energy and matter
- Systems and system models
- Stability and change
- Patterns

Click here to read the full standards.

### Essential questions

- How can changes from reactants to products in chemical reactions be represented to demonstrate conservation of mass?
- How are changes in energy for chemical reaction systems modeled and described?
- How can patterns in chemical bonding and reactivity be used to predict the outcome of a reaction?

## Lesson notes

### Lesson 1: Representing chemical reactions

**PEs: HS-PS1-2**
**DCIs: HS-PS1.B.3**

**Resources**

| Video | Article | Exercise |
|---|---|---|
| 2 | 1 | 2 |

| Objectives | Teaching tips |
|---|---|
| • Interpret the meaning of symbols and numbers in **chemical equations**, and identify the **reactants** and **products**.<br>• Explain the difference in meaning between **subscripts** and **coefficients** in a balanced chemical equation.<br>• Use subscripts and coefficients to determine the number of each kind of atom present in the reactants and products of a balanced chemical equation. | • Introduce a balanced chemical equation and ask students to describe all of the components using a given list of vocabulary words (e.g., reactants, products, coefficient, subscript, solid, liquid, gas, aqueous, atom, molecule, formula unit, yields).<br>• Ask students to translate a few balanced chemical equations into words as complete sentences. This provides practice interpreting the symbols in chemical equations and emphasizes the value of using the equation notation as effective shorthand.<br>• Review and reinforce students' chemical formula writing skills from Unit 3 by asking them to translate reactions from words into chemical equations.<br>• Have students cut out pieces of paper in different colors to represent different types of atoms. Guide students in using these to model groups of atoms in chemical equations. Ask them to model the difference, for example, between 2 NaCl and $Na_2Cl_2$ or $O_2$ and 2 O. Have students use the models to count the number of each type of atom in the reactants and in the products. Ask what they notice about the total number of each kind of atom in the reactants vs the products (they are the same), and note that this will be important to remember in the next lesson.<br>• Once students become comfortable using the physical model to interpret subscripts and coefficients, provide opportunities for them to practice determining the number of each kind of atom present in the reactants and products from the chemical equation alone, without relying on the model. |

### Lesson 2: Balancing chemical equations

**PEs: HS-PS1-2**
**DCIs: HS-PS1.B.3**

**Resources**

| Video | Exercise |
|---|---|
| 4 | 2 |

| Objectives | Teaching tips |
|---|---|
| • Explain why the **law of conservation of mass** requires chemical equations to be balanced with coefficients.<br>• Use coefficients to balance chemical equations. | • Before getting into the mechanics of balancing chemical equations, introduce the law of conservation of mass. Ask students to remember what they noticed in Lesson 1 about the total number of each kind of atom in the reactants versus the products of a chemical equation. Connect conservation of mass to the idea that chemical reactions involve the *rearrangement* of atoms, not their destruction or creation.<br>• Have students work in pairs to analyze an *unbalanced* chemical equation without any coefficients. Ask them to: |

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| | ◦ Determine the number of each kind of atom in the reactants and in the products.<br>◦ Explain how the equation as written violates the law of conservation of mass.<br>◦ Brainstorm one or more strategies for adjusting the equation to be consistent with conservation of mass.<br>Invite students to share and discuss their different approaches with the class. Guide students to recognize why adding coefficients is the appropriate strategy to "balance" a chemical equation, and ask them to explain *why* changing subscripts is *not* an acceptable strategy.<br>• Guide students in using the PhET Balancing Chemical Equations simulation to visualize what is changing and not changing about the components of a chemical equation when they adjust coefficients. Have students work in small groups, using the *Intro* section of the simulation to answer questions like the ones below. Instruct students to click on the *Tools* menu in the upper right of the application screen to select the balance or bar graph tool.<br>◦ When all of the coefficients are 1, what do you notice about the total number of each type of atom in the reactants and the products (represented on the balances or in the bar graphs)?<br>◦ When you increase a reactant's coefficient from 1 to 2, what appears in the reactants box? How do the balances or bar graphs change?<br>◦ How does changing coefficients eventually lead to a balanced chemical equation? When do you know that the equation is balanced?<br>◦ Do the subscripts in the chemical formulas of the reactants or products change during the process of balancing the chemical equation? What evidence do you have from the simulation?<br>• Discuss students' discoveries from the simulation. Emphasize that the process of balancing a chemical equation is simply a way of representing on paper what actually occurs in a reaction. Clarify that adding coefficients to a chemical equation does not add to or change the chemical reaction. Rather, it adjusts the *equation* to correctly reflect the ratios in which the substances react.<br>• Challenge students to practice their equation balancing skills with the *Game* section of the PhET Balancing Chemical Equations simulation.<br>• Implement the hands-on activity "*How can a big log turn into a tiny pile of ash when it burns?*" in order for students to observe conservation of mass experimentally and connect it to conservation of atoms through particle diagrams and balanced chemical equations (see Lesson 3). |

| Lesson 3: Hands-on science activity<br>*How can a big log turn into a tiny pile of ash when it burns?*<br><br>PEs: HS-PS1-7<br>DCIs: HS-PS1.B.1, HS-PS1.B.3 | Resources<br>Activity 1 |
|---|---|
| **Description** | **Links** |
| Students will **carry out an investigation** and **analyze and interpret data** to determine what happens to the amount of matter during a physical or chemical change. They will use their experimental data as evidence to support the claim that mass is conserved. | • Full activity overview (Khan Academy article)<br>• Student activity guide (Doc \| PDF)<br>• Teacher guide (Doc \| PDF) |

| Lesson 4: Energy of chemical reactions<br><br>PEs: HS-PS1-4<br>DCIs: HS-PS1.A.4, HS-PS1.B.1 | Resources<br>Video 1<br>Exercise 2 |
|---|---|
| **Objectives** | **Teaching tips** |
| • Describe what happens in an **exothermic** or **endothermic reaction** in terms of the overall change in energy from reactants to products, the transfer of energy between the reaction system and surroundings, and the resulting change in temperature of the surroundings.<br>• Analyze an **energy diagram** for a chemical reaction to identify the minimum energy required to initiate the reaction, the energy released when the product bonds form, the overall change in energy, and whether the reaction is exothermic or endothermic.<br>• Draw an energy diagram to represent an exothermic or endothermic chemical reaction. | • ***Note:*** *Unit 7 provides a deep dive into thermal energy transfer within and between systems, including calorimetry. Emphasis in this lesson is on the connection between the overall change in energy for a chemical reaction and the energy required to break bonds in the reactants versus energy released when new bonds form in the products.*<br>• Utilize quick demonstrations that allow students to experience exothermic and endothermic processes, such as dissolving calcium chloride in water or activating a hand warmer (exothermic) and dissolving ammonium chloride or an effervescent tablet in water (endothermic). Allow students to touch the outside of the reaction vessel, which will feel "hot" (exothermic) or "cold" (endothermic).<br>• Guide students in drawing a diagram that models an exothermic reaction they observed. As a class, define the system, indicate the direction of energy transfer, and note what happens to the temperature of the surroundings as the reaction occurs. Next, have students work in pairs to model an endothermic reaction that they observed. Introduce the terms "exothermic" and "endothermic," and label the two models.<br>• Show students a balanced chemical equation and an energy diagram for an exothermic reaction, such as the formation of water from hydrogen and oxygen gas.<br>    ○ Label the reactants and products on the energy diagram, then have students use their colored paper models from Lesson 1 or molecular model kits to build the reactant molecules.<br>    ○ Ask students to model breaking the bonds between atoms in the reactants. Explain that this process requires energy, and point out the activation energy in the diagram.<br>    ○ Then, have students model rearranging and bonding the atoms |

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| | ◦ to form product molecules. Explain that this process of bond formation releases energy, and point out how the diagram slopes downward to the final products.<br>◦ Emphasize that the overall change in energy for an exothermic reaction is negative, because more energy is released when bonds form than is required to break bonds.<br>• After guiding students through the energy diagram for an exothermic reaction, ask them to consider how the energy diagram would be similar and different for an endothermic reaction. Have them work in pairs or small groups on whiteboards to sketch and label what they think the diagram should look like, then discuss as a class.<br>• Have students fill in a T-chart comparing exothermic and endothermic reactions in terms of:<br>◦ Energy required for bond breaking vs energy released when new bonds form (more/less)<br>◦ Sign of overall energy change (positive/negative)<br>◦ Direction of energy transfer (into/out of reaction system)<br>◦ Change in temperature of surroundings (increase/decrease)<br>◦ Height of reactants versus products on the *y*-axis of an energy diagram (higher/lower) |

## Lesson 5: Types of chemical reactions

**PEs: HS-PS1-2**
**DCIs: HS-PS1.B.1, HS-PS1.B.3**

**Resources**

| Video | Article | Exercise |
|---|---|---|
| 3 | 1 | 2 |

| Objectives | Teaching tips |
|---|---|
| • Identify a reaction as a **synthesis**, **decomposition**, **combustion**, or **single replacement reaction**, based on how the reactants rearrange to form products.<br>• Explain the concepts of **oxidation** and **reduction** in terms of electron transfer, and identify elements that are oxidized or reduced in a chemical reaction.<br>• Apply knowledge of chemical bonding and reaction patterns to predict the products and write balanced chemical equations for single replacement reactions and simple synthesis and decomposition reactions.<br>• Use the relative reactivities of elements to determine if a given single replacement reaction will occur. | • Provide students with visual and tangible frames of reference for understanding reaction patterns through videos or live demonstrations of exciting chemical reactions. Ensure that appropriate safety precautions are in place for all live demonstrations. Example demonstrations:<br>◦ *Synthesis:* iron and sulfur react to form iron(II) sulfide.<br>◦ *Decomposition:* hydrogen peroxide breaks down into water and oxygen in the presence of a catalyst.<br>◦ *Single replacement:* magnesium reacts with hydrochloric acid to produce aqueous magnesium chloride and hydrogen gas.<br>◦ *Combustion:* methane reacts with oxygen to produce carbon dioxide and water vapor.<br>• Provide students with sets of 12-16 index cards, each with an example of a synthesis, decomposition, single replacement, or combustion reaction on it. Have students work in pairs or small groups to organize the reactions into categories based on patterns that they notice, then ask them to share their categories and reasoning with the class. Use this as a segue to introduce the names of the four types of reactions and their characteristic patterns.<br>• Guide students to diagram redox reactions by having them identify charges on atoms in the reactants and products and use arrows to indicate the number of electrons gained or lost. Introduce mnemonic |

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| | devices, such as “OIL RIG” (**O**xidation **I**s **L**oss, **R**eduction **I**s **G**ain), to help students remember that oxidation occurs when an atom *loses* electrons, and reduction is when an atom *gains* electrons.<br>• Provide scaffolded worksheets that break down into discrete steps the process of writing a balanced chemical equation for a reaction from a given set of reactants. For example, ask students to:<br>○ Predict the type of chemical reaction (synthesis, decomposition, combustion, or single replacement) that will occur between the given reactants and provide reasoning.<br>○ Write correct chemical formulas for the products, based on reaction patterns.<br>○ Use the activity series to determine if a single replacement reaction will occur.<br>○ Write state symbols for the products.<br>○ Balance the chemical equation with coefficients.<br>Before removing this scaffolding, have students generate their own lists of steps and questions to consider when approaching these problems. Ask them to consider why it is important to balance the chemical equation last. |

## Related phenomena

### Example phenomenon

What causes a car or a ship to rust, and how can we prevent this process?

### Background information

*Ship with rust on its hull*

When iron is exposed to moisture and air, it undergoes a redox reaction with water and oxygen to produce hydrated iron(III) oxide, commonly known as rust. This reddish-brown solid forms a layer on the surface of the metal that gradually flakes off, exposing more pure iron to the environment, and allowing further oxidation. Over time, if this process is allowed to continue, the iron deteriorates to the point that it loses its structural integrity.

Although the process takes place through several chemical steps, the overall redox reaction for the rusting of iron can be represented as:

$$4Fe(s) + 3O_2(g) + 2H_2O(l) \rightarrow 2Fe_2O_3 \cdot 2H_2O(s)$$

Iron is oxidized as it loses electrons to form $Fe^{3+}$ ions and oxygen is reduced as it gains electrons to form $O^{2-}$ ions. The rate at which this redox reaction occurs depends on a number of factors, including the availability of water through moisture in the air, rain, or a marine environment, the presence of dissolved salts, and the pH of the aqueous environment.

A number of strategies are used to prevent the formation of rust on important structures, vehicles, and everyday objects made of iron or steel (an alloy of iron and carbon), such as bridges, cars, ships, food cans, and nails. The simplest strategy is to cover the iron with paint or coat it with grease or oil to provide a barrier between the metal and water and oxygen in the environment. Similarly, a thin layer of a *less reactive* metal, such as tin, nickel, or chromium, can be electroplated onto the surface of an iron object, effectively sealing it off from the atmosphere. The less reactive metal gradually forms a thin oxide layer that acts as a coating to further protect the pure metal underneath. Soup is packaged in tin-coated cans, and the shiny appearance of chrome bumpers on cars comes from nickel and chromium plating.

Another approach is to "galvanize" iron by coating it with a layer of zinc, usually by dipping the iron object in a vat of molten zinc. This similarly protects against oxidation by acting as a physical barrier between the iron and moist air, but since zinc is *more reactive* than iron, it acts as a "sacrificial metal." Even if the coating is damaged and part of the iron structure becomes exposed, it will not begin to oxidize until all the surrounding zinc has reacted.

Iron also may be alloyed (mixed) with other metals to decrease its tendency to oxidize. For example, stainless steel is a durable, easy to clean, corrosion resistant alloy containing 11% or more chromium. It is used in applications such as cookware, surgical instruments, appliances, and industrial manufacturing equipment.

Exploring this phenomenon helps students develop and master the following understandings:

- [ ] Chemical reactions can be represented by balanced chemical equations.
- [ ] In a redox reaction, one element is oxidized (loses electrons) and another element is reduced (gains electrons).
- [ ] The relative reactivities of different elements will determine which reactions take place.

**Tips for implementing phenomenon-based learning**

- Ideas to encourage student engagement:
    - Ask students to brainstorm examples they have observed of iron structures or objects rusting. Elicit their ideas about what makes a structure or object more likely to rust. Prompt them to think about environmental conditions that might be relevant.
    - Use a simple demonstration to show that both water and oxygen are required for rusting:
        - Label three test tubes A, B, and C, and place a clean iron nail in each.
        - For tube A, add enough tap water to partially cover the nail, then stopper the tube.
        - For tube B, add enough hot distilled water to entirely cover the nail. Pipette about 1 mL of vegetable oil into the tube to form a layer on top of the water, then stopper the tube.
        - For tube C, add a small scoop of solid anhydrous calcium chloride, which will absorb any water present in the air, then stopper the tube.

      After a few days, the nail in tube A (exposed to water and oxygen) will begin to rust, while the nail in tube B (only exposed to water) and the nail in tube C (only exposed to oxygen) will not. It may be useful to set up this demo a few days before introducing it to students so that the rusting process has time to begin. Students can then observe its progress over the following days.

      As a follow up, set up three new test tubes with nails partially covered in tap water, tap water with salt, and tap water with a few drops of vinegar respectively. Ask students to predict what will happen over the next week as they observe the nails. Guide them to connect their observations to environmental factors that increase the rate of rusting, like the salinity of ocean water and the lower pH of acid rain.

- Give students the unbalanced overall chemical equation for iron rusting, and discuss the meaning of the hydrate notation in the product.

  $$Fe(s) + O_2(g) + H_2O(l) \rightarrow Fe_2O_3 \cdot 2H_2O(s)$$

  Ask students to balance the equation and to provide evidence for why rusting is classified as a redox reaction. Students should identify the elements that are oxidized/reduced and determine the number of electrons lost/gained by each element.
- Based on what they have learned so far about the process of rusting, ask students to work in pairs or small groups to come up with practical ways to prevent it. Have students share their ideas with the class. Discuss which solutions are most practical, and use this as a way to introduce rust prevention strategies that students did not suggest.
- Introduce the methods of electroplating iron with tin, nickel, or chromium and galvanizing with zinc. Have students use the metal activity series to determine whether the coating metals are more or less reactive than iron. Guide students to think about what will happen in each case if the metal coating is scratched to expose the iron surface underneath.
- Refer back to the phenomenon question in [Unit 2: *Why is jewelry made out of gold, silver, and platinum, but not sodium or potassium?*] Ask students if they have ever experienced silver jewelry leaving a black residue on their skin. Explain that silver can react with hydrogen sulfide in the air in a process called "tarnishing." Have students predict the products of the reaction, write a balanced chemical equation, identify the type of reaction (both redox and single replacement), and determine which elements are oxidized/reduced and how many electrons are lost/gained.

- Sample prompts to elicit student ideas and encourage discussion:
  - What reactants are required to form rust?
  - In the process of rusting, what element is oxidized and what element is reduced?
  - Why is the rusting of iron structures and objects a problem that needs to be addressed?
  - How does coating iron with paint, grease, oil, or another metal prevent rusting?
  - How does the relative reactivity of zinc versus iron play an important part in preventing galvanized iron from rusting?
  - What are other everyday examples of pure metals undergoing redox reactions with substances in the environment?

### Example phenomenon

How do disposable hand warmers work?

### Background information

Air-activated disposable hand warmers use the rapid oxidation of iron (rusting) to produce enough thermal energy to keep fingers or toes warm for several hours. A typical hand warmer consists of powdered iron, sodium chloride, activated charcoal, and vermiculite enclosed in a gas-permeable pouch. The pouch is stored in an airtight plastic wrapper until ready for use.

*Person holding a disposable hand warmer*

As soon as the wrapper is opened, oxygen and water vapor in the air move through the porous pouch, contacting the iron powder, and starting the redox reaction. Sodium chloride acts as a catalyst to increase the reaction rate; when dissolved in water, the free sodium and chloride ions allow iron to lose electrons and become oxidized more readily. Vermiculite is an absorbent material that retains water needed for the redox reaction and acts as an insulator to keep thermal energy from transferring out of the pouch too quickly. Activated charcoal helps to disperse thermal energy more evenly throughout the pouch and control the rate of the reaction.

The process of rusting (covered in more detail in the example phenomenon above) is exothermic, but it usually occurs relatively slowly, so the thermal energy released is not noticeable. In a hand warmer, the iron oxidizes much more rapidly, releasing thermal energy quickly in a concentrated area, and allowing the temperature of the pouch to reach more than 70°C. The sodium chloride catalyst and the very high surface area of the powdered iron are mainly responsible for the increased reaction rate. Shaking the pouch and blowing on it will also increase the reaction rate, as this brings more oxygen and water into contact with the iron.

Exploring this phenomenon helps students develop and master the following understandings:

- ☐ In a redox reaction, one element is oxidized (loses electrons) and another element is reduced (gains electrons).
- ☐ In an exothermic reaction, the amount of energy released when product bonds form is greater than the amount of energy input required to break reactant bonds, so the overall change in energy is negative.
- ☐ In an exothermic reaction, thermal energy is released from the reaction system to the surroundings, causing the temperature of the surroundings to increase.

**Tips for implementing phenomenon-based learning**

- Divide students into small groups, and provide each group with a packaged hand warmer. Have students open the plastic packaging and take notes on what they observe as they hold and shake the pouch. As a class, discuss students' observations and inferences from their investigations. Prompt student thinking with questions like:
    - Is the reaction in the hand warmer exothermic or endothermic, and what is your evidence?
    - What did you observe when you shook the pouch?
    - Why do you think shaking the pouch had this effect?
    - What do you think is inside the pouch?
- Conduct a simple demonstration to provide students with more information about the contents of the hand warmer pouch:
    - First, cut open a new hand warmer pouch, spread out the contents on a heat-resistant surface, and quickly run a magnet covered in a plastic bag over the powder. The iron will stick to the magnet and separate from the other components.
    - Next, cut open an old hand warmer pouch that has been exposed to air for several days and is no longer releasing thermal energy. Spread out the contents, and run a magnet covered in a plastic bag over the powder. Since most of the iron should have reacted to form iron(III) oxide already, very little material should be attracted to the magnet.

- As a class, discuss students' observations and inferences from the demonstrations. Prompt student thinking with questions like:
  - What evidence do we have that a chemical reaction took place in the hand warmer?
  - Why didn't the reaction start until the pouch was out of the plastic packaging?
  - Why does the hand warmer eventually stop releasing thermal energy, and why can't it be reused?
  - What do you think is inside the pouch before the reaction begins, and what is left after the reaction is over?
- Guide students to recognize that the chemical reaction taking place in the hand warmer pouch is iron (magnetic) being oxidized to form rust (not magnetic) when the pouch is removed from its packaging and exposed to oxygen and water in the air.

- Review the overall balanced chemical equation for iron rusting discussed in the previous example phenomenon, or if you have not yet covered this, give students the unbalanced overall chemical equation, and discuss the meaning of the hydrate notation in the product.

$$Fe(s) \ + \ O_2(g) \ + \ H_2O(l) \rightarrow Fe_2O_3 \cdot 2H_2O(s)$$

  Ask students to balance the equation and to provide evidence for why rusting is classified as a redox reaction. Students should identify the elements that are oxidized/reduced and determine the number of electrons lost/gained by each element.
- Have students work in pairs or small groups on whiteboards to draw an energy diagram for the chemical reaction in the hand warmer. Ask them to label the diagram with reactants, products, energy input to break reactant bonds, energy released in forming product bonds, and overall energy change for the reaction. Ask students to provide evidence from their observations that supports their diagram.
- Although reaction rates are not covered until Unit 9, this phenomenon provides a natural opportunity to introduce the effect of a catalyst and the surface area of a solid reactant on reaction rate. Ask students to consider why the rusting of a car bumper doesn't seem to produce a lot of thermal energy in the same way that the production of rust in a hand warmer does. Guide them to recognize that the energy is released more quickly in the hand warmer reaction.
  - Identify sodium chloride as a "catalyst" in the reaction. Show students how an energy diagram changes when a catalyst is present, emphasizing that the energy input required to break reactant bonds and the energy released in forming product bonds both decrease, but the overall energy change for the reaction remains the same.
  - Note that the iron in the hand warmer is powdered, rather than one solid chunk. Ask students how they think this might speed up the reaction. Demonstrate the effect of surface area by placing a sugar cube and an equal mass of granulated sugar in two separate containers of water. Ask students to predict what will happen when you stir each mixture for 5-10 seconds, then allow them to observe this process. Guide students to understand that interactions between particles only happen at the surface of a solid substance, so the more exposed surface there is, the faster the process will occur. Revisit this example in Unit 9.

## Common student misconceptions

**Possible misconception:** *It is okay to change subscripts in order to "balance" a chemical equation.*

Students may think it is acceptable to reconcile the number of each kind of atom in the reactants and products of a chemical equation by changing the subscripts in chemical formulas. This misconception may arise, in part, because students do not fully understand the different meanings of subscripts and coefficients.

**Critical concepts**

- A subscript indicates the number of atoms of an element that exist within a single unit of a substance.
- If the subscripts in a chemical formula change, then the formula represents an entirely different substance.
- A coefficient indicates how many separate units of a substance are present in the reaction.
- Changing coefficients changes the number of reacting units in a chemical equation, but it does not change the identities of the substances in the reaction.

**How to address this misconception**

Introduce examples like water ($H_2O$) and hydrogen peroxide ($H_2O_2$) to show that chemical formulas with the same kinds of atoms and different subscripts represent different substances with different properties. Have students represent balanced chemical equations with particle diagrams or physical models to demonstrate how coefficients change the number of molecules or formula units involved in the reaction without changing the identities of the reactants or products. Encourage students to explore the PhET Balancing Chemical Equations simulation to reinforce this understanding.

**Possible misconception:** *Pure elements always appear as individual atoms in chemical equations.*

Students often forget about the existence of diatomic elements that naturally form stable bonded pairs. Alternatively, students may think that diatomic elements must *always* have a subscript of 2, even in a compound. Both of these misconceptions may lead to errors in writing balanced chemical equations.

**Critical concepts**

- Certain elements are found as stable diatomic molecules in their pure elemental forms.
- Atoms are most stable when they have full valence shells. Therefore, atoms will bond with other atoms by transferring or sharing electrons to achieve full valence shells (octet rule).

**How to address this misconception**

Address early on that the elements H, N, O, F, Cl, Br, and I, are diatomic in their pure elemental forms (when not bonded with other elements). It can be helpful to note that these elements are all "gens"—hydro**gen**, nitro**gen**, oxy**gen**, and the halo**gens** (fluorine, chlorine, bromine, and iodine). Ask students to draw Lewis diagrams for $H_2$, $O_2$, and $N_2$ and to explain why these elements are more stable as diatomic molecules (review from Unit 3). Next, ask students to draw Lewis diagrams for sodium hydride (NaH), aluminum oxide ($Al_2O_3$), and magnesium nitride ($Mg_3N_2$). Emphasize that diatomic elements will bond with *other* elements in various different ratios that depend on valence electrons and the octet rule, so they will not necessarily have a subscript of 2. Provide example problems where diatomic elements are reactants or products, so that students practice recognizing and writing the correct formulas for diatomic elements.

**Possible misconception:** *Bond formation is an endothermic process, and bond breaking is an exothermic process.*

Students may believe that energy must be used to bond atoms together to form something new, resulting in an endothermic process. Students also may believe that when bonds are broken, the energy stored within the bonds is released, resulting in an exothermic process. These ideas often stem from studying anabolic and catabolic processes in biology, where students learn that anabolic processes “build up” molecules and require energy, while catabolic processes “break down” molecules and release energy.

**Critical concepts**

- Bond breaking requires an input of energy, while bond formation results in a release of energy.
- The overall energy change for a chemical reaction is the difference between the energy required to break reactant bonds and the energy released when product bonds form.
- If the amount of energy released is greater than the amount of energy input required, the reaction is exothermic. Conversely, if the amount of energy released is less than the amount of energy input required, the reaction is endothermic.

**How to address this misconception**

Use energy diagrams and molecular models to help students visualize changes in energy and bonding during chemical reactions. Emphasize that a minimum amount of energy is required to start a reaction (activation energy). The energy diagram shows an initial increase in energy to break bonds, even for an exothermic reaction like combustion of a hydrocarbon. Energy in the reaction system then decreases as new bonds form to make products. Refer to reactions that students observed in class or explored in the phenomena questions. For example, in the combustion of methane, more energy is released from bond formation than is consumed in bond breaking, causing energy to be released from the reaction system. Point out that the overall energy change is represented in the energy diagram as the difference between the height of the reactants and the products on the *y*-axis. If products are lower than reactants, then the overall energy change is negative and the reaction is exothermic. The reverse is true for an endothermic reaction.

# Unit resources

## Student resources

- PhET Balancing Chemical Equations: Use this simulation for support understanding and practicing the process of balancing chemical equations.
- Iron reacts with sulfur: Observe an example of a synthesis reaction in this video.
- Hydrogen peroxide breaks down: Observe an example of a decomposition reaction in this video.
- Magnesium reacts with hydrochloric acid: Observe an example of a single replacement reaction in this video.
- Methane reacts with oxygen: Observe an example of a combustion reaction in this video.
- Unit 3 (Chemical bonding): Use the resources in this Khan Academy unit to review writing chemical formulas for ionic and covalent compounds.
- Unit 7 (Thermochemistry): Use the resources in this Khan Academy unit to learn more about energy changes in chemical reactions.
- Unit 9 (Reaction rates and equilibrium): Use the resources in this Khan Academy unit to learn more about factors that affect chemical reaction rates.
- Article and video note taking template (Doc | PDF): Use this printable template for structured note taking on the articles and videos in this unit.

## Classroom implementation resources

- PubChem Periodic Table of Elements and Printable Periodic Tables: Download various versions of the periodic table, including a blank template.
- Weekly Khan Academy quick planning guide (Doc | PDF): Use this template to easily plan your week.
- Using Khan Academy in the classroom (Doc | PDF): Learn about teaching strategies and structures to support your students in their learning with Khan Academy.
- Differentiation strategies for the classroom (Doc | PDF): Read about strategies to support the learning of all students.
- Using phenomena with the NGSS: Learn more about how to incorporate phenomena into NGSS-aligned lessons.
- Hands-on science activities from Khan Academy: Choose from Khan Academy's collection of high-quality, ready-to-use, and free hands-on science activities. Each one is engaging, three-dimensional, phenomenon-based, and simple to implement.

# NGSS standards reference guide

## Performance expectations

- **HS-PS1-2:** Construct and revise an explanation for the outcome of a simple chemical reaction based on the outermost electron states of atoms, trends in the periodic table, and knowledge of the patterns of chemical properties.
- **HS-PS1-4:** Develop a model to illustrate that the release or absorption of energy from a chemical reaction system depends upon the changes in total bond energy.
- **HS-PS1-7:** Use mathematical representations to support the claim that atoms, and therefore mass, are conserved during a chemical reaction.

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## Disciplinary core ideas

- **HS-PS1.A.4:** A stable molecule has less energy than the same set of atoms separated; one must provide at least this energy in order to take the molecule apart.
- **HS-PS1.B.1:** Chemical processes, their rates, and whether or not energy is stored or released can be understood in terms of the collisions of molecules and the rearrangements of atoms into new molecules, with consequent changes in the sum of all bond energies in the set of molecules that are matched by changes in kinetic energy.
- **HS-PS1.B.3:** The fact that atoms are conserved, together with knowledge of the chemical properties of the elements involved, can be used to describe and predict chemical reactions.

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## Science and engineering practices (SEPs)

- **Developing and using models**: Students progress to using, synthesizing, and developing models to predict and show relationships among variables between systems and their components in the natural and designed worlds.
- **Planning and carrying out investigations:** Students progress to include investigations that provide evidence for and test conceptual, mathematical, physical, and empirical models.
- **Analyzing and interpreting data:** Students progress to introducing more detailed statistical analysis, the comparison of data sets for consistency, and the use of models to generate and analyze data.
- **Engaging in argument from evidence:** Students progress to using appropriate and sufficient evidence and scientific reasoning to defend and critique claims and explanations about the natural and designed world(s). Arguments may also come from current scientific or historical episodes in science.

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## Crosscutting concepts (CCCs) and their implementation

| Crosscutting concept | Unit implementation |
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| **Systems and system models:** Defining the system under study—specifying its boundaries and making explicitly a model of that system—provides tools for understanding and testing ideas that are applicable throughout science and engineering. | Students define systems in order to study the transfer of thermal energy and/or matter between a system and its surroundings. |

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| **Energy and matter** (Flows, cycles, and conservation): Tracking fluxes of energy and matter into, out of, and within systems helps one understand the systems' possibilities and limitations. | Students predict, observe, and quantify the flow of thermal energy and/or matter into, out of, and within systems as matter undergoes chemical reactions. |
| **Stability and change:** For natural and built systems alike, conditions of stability and determinants of rates of change or evolution of a system are critical elements of study. | Students observe changes in matter via physical changes and chemical reactions, and collect evidence that the total amount of matter remains stable. |
| **Patterns:** Observed patterns of forms and events guide organization and classification, and they prompt questions about relationships and the factors that influence them. | Students use patterns in reactivity to recognize common types of chemical reactions and predict the products of reactions. |