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Predicting precipitation reactions worksheet

In order to continue enjoying our site, we ask that you confirm your identity as a human. Thank you very much for your cooperation. In order to continue enjoying our site, we ask that you confirm your identity as a human. Thank you very much for your cooperation. Humans interact with one another in various and complex ways, and we classify these interactions according to common patterns of behavior. When two humans exchange information, we say they are communicating. When they exchange blows with their fists or feet, we say they are fighting. Faced with a wide range of varied interactions between chemical substances, scientists have likewise found it convenient (or even necessary) to classify chemical interactions by identifying common patterns of reactivity. This module will provide an introduction to three of the most prevalent types of chemical reactions: precipitation, acid-base, and oxidation-reduction. A precipitation reaction is one in which dissolved substances react to form one (or more) solid products. Many reactions of this type involve the exchange of ions between ionic compounds in aqueous solution and are sometimes referred to as double displacement, double replacement, or metathesis reactions. These reactions are common in nature and are responsible for the formation of coral reefs in ocean waters and kidney stones in animals. They are used widely in industry for production of a number of commodity and specialty chemicals. Precipitation reactions also play a central role in many chemical analysis techniques, including spot tests used to identify metal ions and gravimetric methods for determining the composition of matter (see the last module of this chapter). The extent to which a substance may be dissolved in water, or any solvent, is quantitatively expressed as its solubility, defined as the maximum concentration of a substance that can be achieved under specified conditions. Substances with relatively large solubilities are said to be soluble. A substance will precipitate when solution conditions are such that its concentration exceeds its solubility. Substances with relatively low solubilities are said to be insoluble, and these are the substances that readily precipitate from solution. More information on these important concepts is provided in the text chapter on solutions. For purposes of predicting the identities of solids formed by precipitation reactions, one may simply refer to patterns of solubility that have been observed for many ionic compounds (Table 8). Soluble compounds contain group 1 metal cations (Li⁺, Na⁺, K⁺, Rb⁺, and Cs⁺) and ammonium ion (NH₄⁺) the halide ions (Cl⁻, Br⁻, and I⁻) the acetate (C₂H₃O₂⁻), bicarbonate (HCO₃⁻), nitrate (NO₃⁻), and chlorate (ClO₃⁻) ions the sulfate (SO₄²⁻) ion Exceptions to these solubility rules include halides of Ag⁺, Hg₂²⁺, and Pb₂⁺ sulfates of Ag⁺, Ba²⁺, Ca²⁺, Hg₂²⁺, Hg₂²⁺, Pb₂⁺, and Sr²⁺ Insoluble compounds contain carbonate (CO₃²⁻), chromate (CrO₄²⁻), phosphate (PO₄³⁻), and sulfide (S²⁻) ions hydroxide ion (OH⁻) Exceptions to these insolubility rules include compounds of these anions with group 1 metal cations and ammonium ion hydroxides of group 1 metal cations and Ba²⁺ Table 8. Solubilities of Common Ionic Compounds in Water A vivid example of precipitation is observed when solutions of potassium iodide and lead nitrate are mixed, resulting in the formation of solid lead iodide: [latex]2\text{KI}(aq) + \text{Pb}(\text{NO}_3)_2(aq) \rightarrow \text{PbI}_2(s) + 2\text{KNO}_3(aq)] This observation is consistent with the solubility guidelines: The only insoluble compound among all those involved is lead iodide, one of the exceptions to the general solubility of iodide salts. The net ionic equation representing this reaction is: [latex]\text{Pb}^{2+}(aq) + 2\text{I}^{-}(aq) \rightarrow \text{PbI}_2(s)] Lead iodide is a bright yellow solid that was formerly used as an artist's pigment known as iodine yellow (Figure 1). The properties of pure PbI₂ crystals make them useful for fabrication of X-ray and gamma ray detectors. Figure 1. A precipitate of PbI₂ forms when solutions containing Pb²⁺ and I⁻ are mixed. (credit: Der Kreole/Wikimedia Commons) The solubility guidelines in Table 9 may be used to predict whether a precipitation reaction will occur when solutions of soluble ionic compounds are mixed together. One merely needs to identify all the ions present in the solution and then consider if possible cation/anion pairing could result in an insoluble compound. For example, mixing solutions of silver nitrate and sodium fluoride will yield a solution containing Ag⁺, NO₃⁻, Na⁺, and F⁻ ions. Aside from the two ionic compounds originally present in the solutions, AgNO₃ and NaF, two additional ionic compounds may be derived from this collection of ions: NaNO₃ and AgF. The solubility guidelines indicate all nitrate salts are soluble but that AgF is one of the exceptions to the general solubility of fluoride salts. A precipitation reaction, therefore, is predicted to occur, as described by the following equations: [latex]\text{AgNO}_3(aq) + \text{NaF}(aq) \rightarrow \text{AgF}(s) + \text{NaNO}_3(aq)] (molecular) [latex]\text{Ag}^{+}(aq) + \text{F}^{-}(aq) \rightarrow \text{AgF}(s)] (net ionic) Predicting Precipitation Reactions Predict the result of mixing reasonably concentrated solutions of the following ionic compounds. If precipitation is expected, write a balanced net ionic equation for the reaction. (a) potassium sulfate and barium nitrate (b) lithium chloride and silver acetate (c) lead nitrate and ammonium carbonate Solution (a) The two possible products for this combination are KNO₃ and BaSO₄. The solubility guidelines indicate BaSO₄ is insoluble, and so a precipitation reaction is expected. The net ionic equation for this reaction, derived in the previous module, is [latex]\text{Ba}^{2+}(aq) + \text{SO}_4^{2-}(aq) \rightarrow \text{BaSO}_4(s)] (b) The two possible products for this combination are LiC₂H₃O₂ and AgCl. The solubility guidelines indicate AgCl is insoluble, and so a precipitation reaction is expected. The net ionic equation for this reaction, derived in the manner detailed in the previous module, is [latex]\text{Ag}^{+}(aq) + \text{Cl}^{-}(aq) \rightarrow \text{AgCl}(s)] (c) The two possible products for this combination are PbCO₃ and NH₄NO₃. The solubility guidelines indicate PbCO₃ is insoluble, and so a precipitation reaction is expected. The net ionic equation for this reaction, derived in the manner detailed in the previous module, is [latex]\text{Pb}^{2+}(aq) + \text{CO}_3^{2-}(aq) \rightarrow \text{PbCO}_3(s)] Check Your Learning Which solution could be used to precipitate the barium ion, Ba²⁺, in a water sample: sodium chloride, sodium hydroxide, or sodium sulfate? What is the formula for the expected precipitate? An acid-base reaction is one in which a hydrogen ion, H⁺, is transferred from one chemical species to another. Such reactions are of central importance to numerous natural and technological processes, ranging from the chemical transformations that take place within cells and the lakes and oceans, to the industrial-scale production of fertilizers, pharmaceuticals, and other substances essential to society. The subject of acid-base chemistry, therefore, is worthy of thorough discussion, and a full chapter is devoted to this topic later in the text. For purposes of this brief introduction, we will consider only the more common types of acid-base reactions that take place in aqueous solutions. In this context, an acid is a substance that will dissolve in water to yield hydronium ions, H₃O⁺. As an example, consider the equation shown here: [latex]\text{HCl}(aq) + \text{H}_2\text{O}(l) \rightarrow \text{H}_3\text{O}^{+}(aq) + \text{Cl}^{-}(aq)] The process represented by this equation confirms that hydrogen chloride is an acid. When dissolved in water, H₃O⁺ ions are produced by a chemical reaction in which H⁺ ions are transferred from HCl molecules to H₂O molecules (Figure 2). Figure 2. When hydrogen chloride gas dissolves in water, (a) it reacts as an acid, transferring protons to water molecules to yield (b) hydronium ions and solvated chloride ions. The nature of HCl is such that its reaction with water as just described is essentially 100% efficient. Virtually every HCl molecule that dissolves in water will undergo this reaction. Acids that completely react in this fashion are called strong acids, and HCl is one among just a handful of common acid compounds that are classified as strong (Table 9). A far greater number of compounds behave as weak acids and only partially react with water, leaving a large majority of dissolved molecules in their original form and generating a relatively small amount of hydronium ions. Weak acids are commonly encountered in nature, being the substances partly responsible for the tangy taste of citrus fruits, the stinging sensation of insect bites, and the unpleasant smells associated with body odor. A familiar example of a weak acid is acetic acid, the main ingredient in food vinegars: [latex]\text{CH}_3\text{CO}_2\text{H}(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{CH}_3\text{CO}_2^{-}(aq) + \text{H}_3\text{O}^{+}(aq)] When dissolved in water under typical conditions, only about 1% of acetic acid molecules are present in the ionized form. [latex]\text{CH}_3\text{CO}_2\text{H}(aq) \rightleftharpoons \text{CH}_3\text{CO}_2^{-}(aq) + \text{H}^{+}(aq)] (a) Fruits such as oranges, lemons, and grapefruit contain the weak acid citric acid. (b) Vinegars contain the weak acid acetic acid. (credit: modification of work by Scott Bauer; credit by: modification of work by Brücke-Osteuropa/Wikimedia Commons) Compound Formula Name in Aqueous Solution HBr hydrobromic acid HCl hydrochloric acid HI hydroiodic acid HNO₃ nitric acid HClO₄ perchloric acid H₂SO₄ sulfuric acid Table 9. Common Strong Acids A base is a substance that will dissolve in water to yield hydroxide ions, OH⁻. The most common bases are ionic compounds composed of alkali or alkaline earth metal cations (groups 1 and 2) combined with the hydroxide ion—for example, NaOH and Ca(OH)₂. When these compounds dissolve in water, hydroxide ions are released directly into the solution. For example, KOH and Ba(OH)₂ dissolve in water and dissociate completely to produce cations (K⁺ and Ba²⁺, respectively) and hydroxide ions, OH⁻. These bases, along with other hydroxides that completely dissociate in water, are considered strong bases. Consider as an example the dissolution of lye (sodium hydroxide) in water: [latex]\text{NaOH}(s) \rightarrow \text{Na}^{+}(aq) + \text{OH}^{-}(aq)] This equation confirms that sodium hydroxide is a base. When dissolved in water, NaOH dissociates to yield Na⁺ and OH⁻ ions. This also true for any other ionic compound containing hydroxide ions. Since the dissociation process is essentially complete when ionic compounds dissolve in water under typical conditions, NaOH and other ionic hydroxides are all classified as strong bases. Unlike ionic hydroxides, some compounds produce hydroxide ions when dissolved by chemically reacting with water molecules. In all cases, these compounds react only partially and so are classified as weak bases. These types of compounds are also abundant in nature and important commodities in various technologies. For example, global production of the weak base ammonia is typically well over 100 metric tons annually, being widely used as an agricultural fertilizer, a raw material for chemical synthesis of other compounds, and an active ingredient in household cleaners (Figure 4). When dissolved in water, ammonia reacts partially to yield hydroxide ions, as shown here: [latex]\text{NH}_3(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{NH}_4^{+}(aq) + \text{OH}^{-}(aq)] This is, by definition, an acid-base reaction, in this case involving the transfer of H⁺ ions from water molecules to ammonia molecules. Under typical conditions, only about 1% of the dissolved ammonia is present as NH₄⁺ ions. Figure 4. Ammonia is a weak base used in a variety of applications. (a) Pure ammonia is commonly applied as an agricultural fertilizer. (b) Dilute solutions of ammonia are effective household cleansers. (credit: modification of work by National Resources Conservation Service; credit by: modification of work by pat00139) The chemical reactions described in which acids and bases dissolved in water produce hydronium and hydroxide ions, respectively, are, by definition, acid-base reactions. In these reactions, water serves as both a solvent and a reactant. A neutralization reaction is a specific type of acid-base reaction in which the reactants are an acid and a base; the products are often a salt and water, and neither reactant is the water itself: [latex]\text{acid} + \text{base} \rightarrow \text{salt} + \text{water}] To illustrate a neutralization reaction, consider what happens when a typical antacid such as milk of magnesia (an aqueous suspension of solid Mg(OH)₂) is ingested to ease symptoms associated with excess stomach acid (HCl): [latex]\text{Mg}(\text{OH})_2(s) + 2\text{HCl}(aq) \rightarrow \text{MgCl}_2(aq) + 2\text{H}_2\text{O}(l)] Note that in addition to water, this reaction produces a salt, magnesium chloride. Writing Equations for Acid-Base Reactions Write balanced chemical equations for the acid-base reactions described here. (a) the weak acid hydrogen hypochlorite reacts with water (b) a solution of barium hydroxide is neutralized with a solution of nitric acid Solution (a) The two reactants are provided, HOCl and H₂O. Since the substance is reported to be an acid, its reaction with water will involve the transfer of H⁺ from HOCl to H₂O to generate hydronium ions, H₃O⁺, and hypochlorite ions, OCl⁻. [latex]\text{HOCl}(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{H}_3\text{O}^{+}(aq) + \text{OCl}^{-}(aq)] (b) A double-arrow is appropriate in this equation because it indicates the HOCl is a weak acid that has not reacted completely. (b) The two reactants are provided, Ba(OH)₂ and HNO₃. Since this is a neutralization reaction, the two products will be water and a salt composed of the cation of the ionic hydroxide (Ba²⁺) and the anion generated when the acid transfers its hydrogen ion (NO₃⁻). [latex]\text{Ba}(\text{OH})_2(aq) + 2\text{HNO}_3(aq) \rightarrow \text{Ba}(\text{NO}_3)_2(aq) + 2\text{H}_2\text{O}(l)] Check Your Learning Write the net ionic equation representing the neutralization of any strong acid with an ionic hydroxide. (Hint: Consider the ions produced when a strong acid is dissolved in water.) [latex]\text{H}^{+}(aq) + \text{OH}^{-}(aq) \rightarrow \text{H}_2\text{O}(l)] Explore the microscopic view of strong and weak acids and bases. Earth's atmosphere contains about 20% molecular oxygen, O₂, a chemically reactive gas that plays an essential role in the metabolism of aerobic organisms and in many environmental processes that shape our world. The term oxidation is typically used to describe chemical reactions involving O₂, but its meaning has evolved to refer to a broad and important reaction class known as oxidation-reduction (redox) reactions. A few examples of such reactions will be used to develop a clear picture of this classification. Some redox reactions involve the transfer of electrons between reactant species to yield ionic products, such as the reaction between sodium and chlorine to yield sodium chloride: [latex]2\text{Na}(s) + \text{Cl}_2(g) \rightarrow 2\text{NaCl}(s)] It is helpful to view the process with regard to each individual reactant, that is, to represent the fate of each reactant in the form of an equation called a half-reaction: [latex]2\text{Na}(s) \rightarrow 2\text{Na}^{+}(aq) + 2\text{e}^{-} \text{Cl}_2(g) + 2\text{e}^{-} \rightarrow 2\text{Cl}^{-}(aq)] These equations show that Na atoms lose electrons while Cl atoms (in the Cl₂ molecule) gain electrons, the "s" subscripts for the resulting ions signifying they are present in the form of a solid ionic compound. For redox reactions of this sort, the loss and gain of electrons define the complementary processes that occur: [latex]\text{oxidation} + \text{reduction} = 0] (where oxidation is the loss of electrons and reduction is the gain of electrons) In this reaction, then, sodium is oxidized and chlorine undergoes reduction. Viewed from a more active perspective, sodium functions as a reducing agent (reductant), since it provides electrons to (or reduces) chlorine, which functions as an oxidizing agent (oxidant), as it effectively removes electrons from (oxidizes) sodium. [latex]\text{oxidation} + \text{reduction} = 0] Some redox processes, however, do not involve the transfer of electrons. Consider, for example, a reaction similar to the one yielding NaCl: [latex]2\text{H}_2(g) + \text{Cl}_2(g) \rightarrow 2\text{HCl}(g)] The product of this reaction is a covalent compound, so transfer of electrons in the explicit sense is not involved. To clarify the similarity of this reaction to the previous one and permit an unambiguous definition of redox reactions, a property called oxidation number has been defined. The oxidation number (or oxidation state) of an element in a compound is the charge its atoms would possess if the compound were ionic. The following guidelines are used to assign oxidation numbers to each element in a molecule or ion. The oxidation number of an atom in an elemental substance is zero. The oxidation number of a monatomic ion is equal to the ion's charge. Oxidation numbers for common nonmetals are usually assigned as follows: Hydrogen: +1 (when combined with nonmetals, -1 when combined with metals); Oxygen: -2 in most compounds, sometimes -1 (so-called peroxides, O₂²⁻), very rarely [latex>\frac{1}{2}(\text{O}_2)] (so-called superoxides, O₂⁻), positive values when combined with F (values vary); Halogens: -1 for F always, -1 for other halogens except when combined with oxygen or other halogens (positive oxidation numbers in these cases, varying values). The sum of oxidation numbers for all atoms in a molecule or polyatomic ion equals the charge on the molecule or ion. Note: The proper convention for reporting charge is to write the number first, followed by the sign (e.g., 2+), while oxidation number is written with the reversed sequence, sign followed by number (e.g., +2). This convention aims to emphasize the distinction between these two related properties. Assigning Oxidation Numbers Follow the guidelines in this section of the text to assign oxidation numbers to all the elements in the following species: (a) H₂S (b) SO₃²⁻ (c) Na₂SO₄ Solution (a) According to guideline 1, the oxidation number for H is +1. Using this oxidation number and the compound's formula, guideline 4 may then be used to calculate the oxidation number for sulfur: [latex]2(+1) + x = 0 \Rightarrow x = -2] (b) Guideline 3 suggests the oxidation number for oxygen is -2. Using this oxidation number and the ion's formula, guideline 4 may then be used to calculate the oxidation number for sulfur: [latex]x + 6(-2) = -2 \Rightarrow x = +6] (c) Guideline 3 suggests the oxidation number for sodium is +1. Assuming the usual oxidation number for oxygen (-2 per guideline 3), the oxidation number for sulfur is calculated as directed by guideline 4: [latex]2(+1) + x + 4(-2) = -2 \Rightarrow x = +6] Check Your Learning Assign oxidation states to the elements whose atoms are underlined in each of the following compounds or ions. (a) KNO₃ (b) AlH₃ (c) NH₄⁺ (d) H₂PO₄⁻ (e) N₂O₅ (f) Al³⁺ (g) N₂O₃ (h) P₄ (i) H₂O₂ (j) Br₂ Using the oxidation number concept, an all-inclusive definition of redox reaction has been established. Oxidation-reduction (redox) reactions are those in which one or more elements involved undergo a change in oxidation number. (While the vast majority of redox reactions involve changes in oxidation number for two or more elements, a few interesting exceptions to this rule do exist. Example 4.) Definitions for the complementary processes of this reaction class are correspondingly revised as shown here: [latex]\text{oxidation} = \text{increase in oxidation number} \text{reduction} = \text{decrease in oxidation number} \text{Returning to the reactions used to introduce this topic, they may now both be identified as redox processes. In the reaction between sodium and chlorine to yield sodium chloride, sodium is oxidized (its oxidation number increases from 0 in Na to +1 in NaCl) and chlorine is reduced (its oxidation number decreases from 0 in Cl₂ to -1 in NaCl). In the reaction between molecular hydrogen and chlorine, hydrogen is oxidized (its oxidation number increases from 0 in H₂ to +1 in HCl) and chlorine is reduced (its oxidation number decreases from 0 in Cl₂ to -1 in HCl). Several subclasses of redox reactions are recognized, including combustion reactions in which the reductant (also called a fuel) and oxidant (often, but not necessarily, molecular oxygen) react vigorously and produce significant amounts of heat, and often light, in the form of a flame. Solid rocket-fuel reactions such as the one depicted in Figure 1 in Chapter 4 Introduction are combustion processes. A typical propellant reaction in which solid aluminum is oxidized by ammonium perchlorate is represented by this equation: [latex]10\text{Al}(s) + 6\text{NH}_4\text{ClO}_4(s) \rightarrow 10\text{Al}_2\text{O}_3(s) + 2\text{N}_2(g) + 12\text{H}_2\text{O}(g) + 3\text{Cl}_2(g) + 2\text{HCl}(g)] Watch a brief video showing the test firing of a small-scale, prototype, hybrid rocket engine planned for use in the new Space Launch System being developed by NASA. The first engines firing at 3 s during the engine flame use a liquid fuel/oxidant mixture, and the second, more powerful engines firing at 4 s (yellow flames) use a solid mixture. Single-displacement (replacement) reactions are redox reactions in which an ion in solution is displaced (or replaced) via the oxidation of a metallic element. One common example of this type of reaction is the acid oxidation of certain metals: [latex]\text{Zn}(s) + 2\text{HCl}(aq) \rightarrow \text{ZnCl}_2(aq) + \text{H}_2(g)] Metallic elements may also be oxidized by solutions of other metal salts; for example: [latex]\text{Cu}(s) + 2\text{AgNO}_3(aq) \rightarrow \text{Cu}(\text{NO}_3)_2(aq) + 2\text{Ag}(s)] This reaction may be observed by placing copper wire in a solution containing a dissolved silver salt. Silver ions in solution are reduced to elemental silver at the surface of the copper wire, and the resulting Cu²⁺ ions dissolve in the solution to yield a characteristic blue color (Figure 5). Figure 5. (a) A copper wire is shown next to a solution containing silver(I) ions. (b) Displacement of dissolved silver ions by copper ions results in (c) accumulation of gray-colored silver metal on the wire and development of a blue color in the solution, due to dissolved copper ions. (credit: modification of work by Mark Ott) Describing Redox Reactions Identify which equations represent redox reactions, providing a name for the reaction if appropriate. For those reactions identified as redox, name the oxidant and reductant. (a) [latex]\text{ZnCO}_3(s) \rightarrow \text{ZnO}(s) + \text{CO}_2(g)] (b) [latex]2\text{Fe}(s) + 3\text{Cl}_2(g) \rightarrow 2\text{FeCl}_3(s)] (c) [latex]2\text{H}_2(g) + \text{O}_2(g) \rightarrow 2\text{H}_2\text{O}(g)] (d) [latex]\text{BaCl}_2(aq) + \text{K}_2\text{SO}_4(aq) \rightarrow \text{BaSO}_4(s) + 2\text{KCl}(aq)] (e) [latex]\text{C}_2\text{H}_6(g) + 7\text{O}_2(g) \rightarrow 2\text{CO}_2(g) + 3\text{H}_2\text{O}(g)] (f) [latex]\text{C}_2\text{H}_6(g) + 7\text{O}_2(g) \rightarrow 2\text{CO}_2(g) + 3\text{H}_2\text{O}(g)] (g) [latex]\text{H}_2\text{O}_2(aq) + 2\text{H}^{+}(aq) + 2\text{e}^{-} \rightarrow 2\text{H}_2\text{O}(l)] (h) [latex]\text{H}_2\text{O}_2(aq) + 2\text{H}^{+}(aq) + 2\text{e}^{-} \rightarrow 2\text{H}_2\text{O}(l)] (i) [latex]\text{H}_2\text{O}_2(aq) + 2\text{H}^{+}(aq) + 2\text{e}^{-} \rightarrow 2\text{H}_2\text{O}(l)] (j) [latex]\text{H}_2\text{O}_2(aq) + 2\text{H}^{+}(aq) + 2\text{e}^{-} \rightarrow 2\text{H}_2\text{O}(l)] (k) [latex]\text{H}_2\text{O}_2(aq) + 2\text{H}^{+}(aq) + 2\text{e}^{-} \rightarrow 2\text{H}_2\text{O}(l)] (l) [latex]\text{H}_2\text{O}_2(aq) + 2\text{H}^{+}(aq) + 2\text{e}^{-} \rightarrow 2\text{H}_2\text{O}(l)] (m) [latex]\text{H}_2\text{O}_2(aq) + 2\text{H}^{+}(aq) + 2\text{e}^{-} \rightarrow 2\text{H}_2\text{O}(l)] (n) [latex]\text{H}_2\text{O}_2(aq) + 2\text{H}^{+}(aq) + 2\text{e}^{-} \rightarrow 2\text{H}_2\text{O}(l)] (o) [latex]\text{H}_2\text{O}_2(aq) + 2\text{H}^{+}(aq) + 2\text{e}^{-} \rightarrow 2\text{H}_2\text{O}(l)] (p) [latex]\text{H}_2\text{O}_2(aq) + 2\text{H}^{+}(aq) + 2\text{e}^{-} \rightarrow 2\text{H}_2\text{O}(l)] (q) [latex]\text{H}_2\text{O}_2(aq) + 2\text{H}^{+}(aq) + 2\text{e}^{-} \rightarrow 2\text{H}_2\text{O}(l)] (r) [latex]\text{H}_2\text{O}_2(aq) + 2\text{H}^{+}(aq) + 2\text{e}^{-} \rightarrow 2\text{H}_2\text{O}(l)] (s) [latex]\text{H}_2\text{O}_2(aq) + 2\text{H}^{+}(aq) + 2\text{e}^{-} \rightarrow 2\text{H}_2\text{O}(l)] (t) [latex]\text{H}_2\text{O}_2(aq) + 2\text{H}^{+}(aq) + 2\text{e}^{-} \rightarrow 2\text{H}_2\text{O}(l)] (u) [latex]\text{H}_2\text{O}_2(aq) + 2\text{H}^{+}(aq) + 2\text{e}^{-} \rightarrow 2\text{H}_2\text{O}(l)] (v) [latex]\text{H}_2\text{O}_2(aq) + 2\text{H}^{+}(aq) + 2\text{e}^{-} \rightarrow 2\text{H}_2\text{O}(l)] (w) [latex]\text{H}_2\text{O}_2(aq) + 2\text{H}^{+}(aq) + 2\text{e}^{-} \rightarrow 2\text{H}_2\text{O}(l)] (x) [latex]\text{H}_2\text{O}_2(aq) + 2\text{H}^{+}(aq) + 2\text{e}^{-} \rightarrow 2\text{H}_2\text{O}(l)] (y) [latex]\text{H}_2\text{O}_2(aq) + 2\text{H}^{+}(aq) + 2\text{e}^{-} \rightarrow 2\text{H}_2\text{O}(l)] (z) [latex]\text{H}_2\text{O}_2(aq) + 2\text{H}^{+}(aq) + 2\text{e}^{-} \rightarrow 2\text{H}_2\text{O}(l)] (aa) [latex]\text{H}_2\text{O}_2(aq) + 2\text{H}^{+}(aq) + 2\text{e}^{-} \rightarrow 2\text{H}_2\text{O}(l)] (ab) [latex]\text{H}_2\text{O}_2(aq) + 2\text{H}^{+}(aq) + 2\text{e}^{-} \rightarrow 2\text{H}_2\text{O}(l)] (ac) [latex]\text{H}_2\text{O}_2(aq) + 2\text{H}^{+}(aq) + 2\text{e}^{-} \rightarrow 2\text{H}_2\text{O}(l)] (ad) [latex]\text{H}_2\text{O}_2(aq) + 2\text{H}^{+}(aq) + 2\text{e}^{-} \rightarrow 2\text{H}_2\text{O}(l)] (ae) [latex]\text{H}_2\text{O}_2(aq) + 2\text{H}^{+}(aq) + 2\text{e}^{-} \rightarrow 2\text{H}_2\text{O}(l)] (af) [latex]\text{H}_2\text{O}_2(aq) + 2\text{H}^{+}(aq) + 2\text{e}^{-} \rightarrow 2\text{H}_2\text{O}(l)] (ag) [latex]\text{H}_2\text{O}_2(aq) + 2\text{H}^{+}(aq) + 2\text{e}^{-} \rightarrow 2\text{H}_2\text{O}(l)] (ah) [latex]\text{H}_2\text{O}_2(aq) + 2\text{H}^{+}(aq) + 2\text{e}^{-} \rightarrow 2\text{H}_2\text{O}(l)] (ai) [latex]\text{H}_2\text{O}_2(aq) + 2\text{H}^{+}(aq) + 2\text{e}^{-} \rightarrow 2\text{H}_2\text{O}(l)] (aj) [latex]\text{H}_2\text{O}_2(aq) + 2\text{H}^{+}(aq) + 2\text{e}^{-} \rightarrow 2\text{H}_2\text{O}(l)] (ak) [latex]\text{H}_2\text{O}_2(aq) + 2\text{H}^{+}(aq) + 2\text{e}^{-} \rightarrow 2\text{H}_2\text{O}(l)] (al) 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