Covalent bonds worksheet answers

The number of bonds that each element is able to form is usually equal to the number of unpaired electrons. In order to form a covalent bond, each element has to share one unpaired electron. Fig. 2.29 gives an example of how to make a Lewis dot structure. First, determine how many atoms of each element are needed to satisfy the octet rule for each atom. In the formation of water, an oxygen atom has two unpaired electrons, and each hydrogen atom has one (Fig. 2.29 A). To fill its valence shell, oxygen needs two additional electrons, and hydrogen needs one. One oxygen atom can share its unpaired electrons with two hydrogen atoms, each of which need only one additional electron. The single electrons match up to make pairs (Fig. 2.29 B). The oxygen atom forms two bonds, one with each of two hydrogen atoms; therefore, the formula for water is H2O. When an electron, or dot, from one element is paired with an electron, or dot, from another element, this makes a bond, which is represented by a line (Fig. 2.29 C). The number of bonds that an element can form is determined by the number of electrons in its valence shell (Fig. 2.29.1). Similarly, the number of electrons in the valence shell also determines ion formation. The octet rule applies for covalent bonding, with a total of eight electrons the most desirable number of unshared or shared electrons in the outer valence shell. For example, carbon has an atomic number of six, with two electrons in shell 1 and four electrons in shell 2, its valence shell (see Fig. 2.29.1). This means that carbon needs four electrons to achieve an octet. Carbon is represented with four unpaired electrons (see Fig. 2.29.1). If carbon can share four electrons with other atoms, its valence shell will be full. Most elements involved in covalent bonding need eight electrons to have a complete valence shell. One notable exception is hydrogen (H). Hydrogen can be considered to be in Group 1 or Group 17 because it has properties similar to both groups. Hydrogen can participate in both ionic and covalent bonding. When participating in covalent bonding, hydrogen only needs two electrons to have a full valence shell. As it has only one electron to start with, it can only make one bond. Single Bonds Hydrogen is shown in Fig 2.28 with one electron. In the formation of a covalent hydrogen molecule, therefore, each hydrogen atom forms a single bond, producing a molecule with the formula H2. A single bond is defined as one covalent bond, or two shared electrons, between two atoms. A molecule can have multiple single bonds. For example, water, H2O, has two single bonds, one between each hydrogen atom and the oxygen atom (Fig. 2.29). Figure 2.30 A has additional examples of single bonds. Double Bonds Sometimes two covalent bonds are formed between two atoms by each atom sharing two electrons, for a total of four shared electrons. For example, in the formation of the oxygen molecule, each atom of oxygen forms two bonds to the other oxygen atom, producing the molecule O2. Similarly, in carbon dioxide (CO2), two double bonds are formed between the carbon and each of the two oxygen atoms (Fig. 2.30 B). Triple Bonds In some cases, three covalent bonds can be formed between two atoms. The most common gas in the atmosphere, nitrogen, is made of two nitrogen atoms bonded by a triple bond. Each nitrogen atom is able to share three electrons for a total of six shared electrons in the N2 molecule (Fig. 2.30 C). Polyatomic Ions In addition to elemental ions, there are polyatomic ions. Polyatomic ions are ions that are made up of two or more atoms held together by covalent bonds. Polyatomic ions can join with other polyatomic ions or elemental ions to form ionic compounds. It is not easy to predict the name or charge of a polyatomic ion by looking at the formula. Polyatomic ions found in seawater are given in Table 2.10. Polyatomic ions bond with other ions in the same way that elemental ions bond, with electrostatic forces caused by oppositely charged ions holding the ions together in an ionic compound bond. Charges must still be balanced. Table 2.10. Common polyatomic ions found in seawater Polyatomic Ion Ion Name NH4+ ammonium CO32- carbonate HCO3- bicarbonate NO2- nitrite NO3- nitrate OH- hydroxide PO43- phosphate HPO42- hydrogen phosphate SiO32- silicate SO32- sulfite SO42- sulfate HSO3- bisulfite Fig. 2.31 shows how ionic compounds form from elemental ions and polyatomic ions. For example, in Fig. 2.31 A, it takes two K+ ions to balance the charge of one (SiO2)2- ion to form potassium silicate. In Figure 2.31 B, ammonium and nitrate ions have equal and opposite charges, so it takes one of each to form ammonium nitrate. Polyatomic ions can bond with monatomic ions or with other polyatomic ions to form compounds. In order to form neutral compounds, the total charges must be balanced. Comparison of Ionic and Covalent Bonds A molecule or compound is made when two or more atoms form a chemical bond that links them together. As we have seen, there are two types of bonds: ionic bonds and covalent bonds. In an ionic bond, the atoms are bound together by the electrostatic forces in the attraction between ions of opposite charge. Ionic bonds usually occur between metal and nonmetal ions. For example, sodium (Na), a metal, and chloride (Cl), a nonmetal, form an ionic bond to make NaCl. In a covalent bond, the atoms bond by sharing electrons. Covalent bonds usually occur between nonmetals. For example, in water (H2O) each hydrogen (H) and oxygen (O) share a pair of electrons to make a molecule of two hydrogen atoms single bonded to a single oxygen atom. In general, ionic bonds occur between elements that are far apart on the periodic table. Covalent bonds occur between elements that are close together on the periodic table. Ionic compounds tend to be brittle in their solid form and have very high melting temperatures. Covalent compounds tend to be soft, and have relatively low melting and boiling points. Water, a liquid composed of covalently bonded molecules, can also be used as a test substance for other ionic and covalently compounds. Ionic compounds tend to dissolve in water (e.g., sodium chloride, NaCl); covalent compounds sometimes dissolve well in water (e.g., hydrogen chloride, HCl), and sometimes do not (e.g., butane, C4H10). Properties of ionic and covalent compounds are listed in Table 2.11. Table 2.11. Properties of ionic and covalent compounds Property Ionic Covalent How bond is made Transfer of e- Sharing of e- Bond is between Metals and nonmetals Nonmetals Position on periodic table Opposite sides Close together Dissolve in water? Yes Varies Consistency Brittle Soft Melting temperature High Low The properties listed in Table 2.11 are exemplified by sodium chloride (NaCl) and chlorine gas (Cl2). Like other ionic compounds, sodium chloride (Fig. 2.32 A) contains a metal ion (sodium) and a nonmetal ion (chloride), is brittle, and has a high melting temperature. Chlorine gas (Fig. 2.32 B) is similar to other covalent compounds in that it is a nonmetal and has a very low melting temperature. Dissolving, Dissociating, and Diffusing Ionic and covalent compounds also differ in what happens when they are placed in water, a common solvent. For example, when a crystal of sodium chloride is put into water, it may seem as though the crystal simply disappears. Three things are actually happening. A large crystal (Fig. 2.33 A) will dissolve, or break down into smaller and smaller pieces, until the pieces are too small to see (Fig. 2.33 B). At the same time, the ionic solid dissociates, or separates into its charged ions (Fig 2.33 C). Finally, the dissociated ions diffuse, or mix, throughout the water (Fig 2.34). Ionic compounds like sodium chloride dissolve, dissociate, and diffuse. Covalent compounds, like sugar and food coloring, can dissolve and diffuse, but they do not dissociate. Fig. 2.34, is a time series of drops of food coloring diffusing in water. Without stirring, the food coloring will mix into the water through only the movement of the water and food coloring molecules. Dissociated sodium (Na+) and chloride (Cl-) ions in salt solutions can form new salt crystals (NaCl) as they become more concentrated in the solution. As water evaporates, the salt solution becomes more and more concentrated. Eventually, there is not enough water left to keep the sodium and chloride ions from interacting and joining together, so salt crystals form. This occurs naturally in places like salt evaporation ponds (Fig. 2.35 A), in coastal tidepools, or in hot landlocked areas (Fig. 2.35 B). Salt crystals can also be formed by evaporating seawater in a shallow dish, as in the Recovering Salts from Seawater Activity. This lesson includes 17 additional questions for subscribers. The number of bonds that each element is able to form is usually equal to the number of unpaired electrons. In order to form a covalent bond, each element has to share one unpaired electron. Fig. 2.29 gives an example of how to make a Lewis dot structure. First, determine how many atoms of each element are needed to satisfy the octet rule for each atom. In the formation of water, an oxygen atom has two unpaired electrons, and each hydrogen atom has one (Fig. 2.29 A). 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Salt crystals can also be formed by evaporating seawater in a shallow dish, as in the Recovering Salts from Seawater Activity. ionic and covalent bonds worksheet answers. all about covalent bonds worksheet answers. bonding basics covalent bonds worksheet answers. ionic vs covalent bonds worksheet answers. drawing covalent bonds worksheet answers. chemical compounds ionic and covalent bonds worksheet answers. identifying ionic and covalent bonds worksheet answers. unit 3 covalent bonds worksheet answers