

LESSON CHEM 3.2 – ANALYZING THE PROPERTIES OF COMPOUNDS AND SOLUTIONS

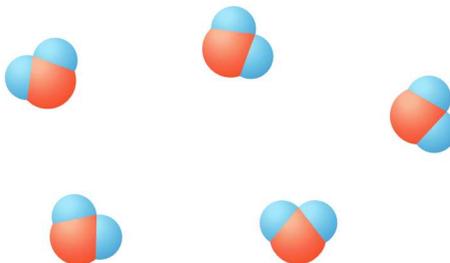
EVIDENCE NOTEBOOK

KEY IDEAS—REVIEW CHEM 3.1

Review the 3.1 lesson again to answer the following questions:

1. Explain AND illustrate the differences in electron distribution between the following types of bonds:
 - a. Covalent
 - b. Polar covalent
 - c. Ionic
2. Dipole moment
 - a. How would you determine where to place a dipole moment on a molecule?
 - b. Illustrate a dipole moment on a *water* molecule:
 - c. Illustrate a dipole moment on an *ammonia* molecule:

3. Hydrogen Bonds: Label the partial negative and partial positive charges on the water molecules. Red spheres indicate oxygen atoms, blue spheres hydrogen atoms. Then, draw dotted lines to represent hydrogen bonds



4. List the different types of bonds in order from strongest to weakest:

BOND TYPE	INTRAMOLECULAR OR INTERMOLECULAR	NOTES	EXAMPLE ILLUSTRATION

CHECKPOINTS—REVIEW CHEM 3.1

5. Water, H_2O , and hydrogen sulfide, H_2S , are both molecules with a bent structure, but the melting point of H_2O is much higher than the melting point of H_2S . Why is this?
 - a. The intermolecular forces in H_2S are much greater than those in H_2O .
 - b. H_2O molecules have hydrogen bonds between them.
 - c. The three-dimensional arrangement of H_2O makes it a nonpolar molecule.
 - d. H_2S has a lower molar mass than H_2O .
6. Which statement best explains why oil and water do not mix?
 - a. Water molecules are polar and oil molecules are nonpolar.
 - b. Oil molecules are polar and water molecules are nonpolar.
 - c. Water molecules are carbon-based and oil molecules are not.
 - d. Oil and water both have covalent bonds and repel one another.
7. Explain whether HCl or HF shows the stronger intermolecular forces and how this is related to trends in the periodic table. Then, explain which of these compounds would have a higher boiling point based on differences in intermolecular forces.

8. What results when equal but opposite charges are present in two regions of a polar molecule?
 - a. dipole
 - b. ionic bond
 - c. electron sea
 - d. crystal lattice
9. Of the three molecules O_2 , HCl , and F_2 , which one has the potential to be polar?
 - a. F_2
 - b. O_2
 - c. HCl
 - d. They all have the potential to be polar.
10. Which reason explains why the boiling point of water (H_2O) is higher than the boiling point of hydrogen sulfide (H_2S)?
 - a. ionic bonding
 - b. covalent bonding
 - c. hydrogen bonding
 - d. London dispersion forces

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KEY IDEAS

11. Avagadro's number is also known as a _____. In scientific notation it is written as _____ . Atomic mass is also termed molar mass. The units for atomic mass (molar mass) are _____ .
12. Convert 5.50 moles of iron into grams of iron. Show your work.
13. Convert 32.4 grams of iron into moles of iron. Show your work.
14. Convert 67.2 grams of iron (II) oxide (FeO) into moles of iron (II) oxide. Show your work.
15. Grams and moles conversions (*show your work*):
 - a. Convert 9.62 g CH₄ into moles.
 - b. Convert 0.0431 mol NaOH into grams.

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16. Match the following:

- | | | |
|-------------|-------|---|
| 1. Solvent | _____ | a. The substance that does the dissolving in a liquid mixture |
| 2. Solute | _____ | b. The liquid mixture |
| 3. Solution | _____ | c. The substance that is dissolved in the liquid mixture. |

17. What is molarity?

18. What is the molarity of a 2.0 liter solution with 3.0 moles of HCl in it?

19. Molarity Math tips (free form personal notes):

20. Convert 125 grams of H₂O (18.0 g/mol) into moles of water. Show your work.

21. Convert 1.68 L of water into mL. Show your work.

22. How many moles of salt are in 458 L container of 4.29 M salt water? Show your work.

23. Molarity practice (*show your work*):

- a. 2.0 mols NaCl of a 3.0 liter solution. What is the molarity?

- b. What is the molarity of 1.23 mols of KCl in a 3.2 liter solution?

- c. If you have a 2.1 molar solution of LiF, how many mols would be in a 1.5 liter sample?

- d. How many liters of a 0.48 M KBr solution would you need to measure out to obtain 0.032 mols of KBr?

- e. A 3.2 liter, 2.1 molar solution of HCl contains how many grams of HCl?

- f. A 2.5 molar solution contains 0.97 g of NaOH. What would be the volume of this solution in mL?

24. What does it mean to have a polar molecule?

25. Structure of water

- a. Draw the Lewis structure for water. Add the dipole moment to your illustration.

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26. Explain how water behaves near an electric charge. Why does it do this?

27. Illustrate how water positions itself to dissolve salt, NaCl.

28. Explain the following terms:

a. Unsaturated solution

b. Saturated solution

c. Supersaturated solution

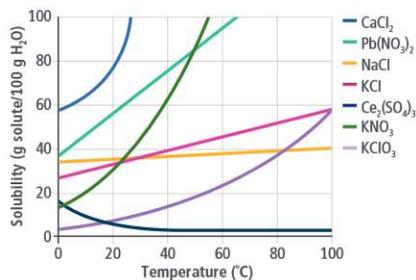
29. Electrical conduction

a. What is required in a substance in order for it to conduct electricity?

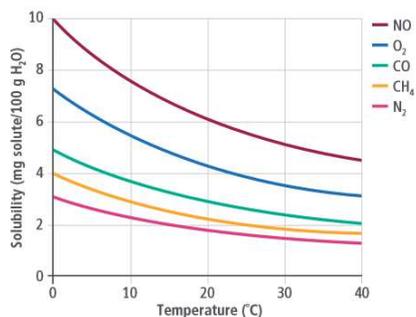
b. An electrolyte is a substance that when dissolved in a solvent, increases the solvent's conductivity. An example of electrolytes include:

30. Factors that affect solubility

- a. Look at the graph below and determine the trend of how temperature affects the solubility of *solutes*.



- b. Look at the graph below and determine the trend of how temperature affects the solubility of *gasses*



31. How is the solubility of gasses in solution related to the external pressure? This is explained by what law?

32. What is molality?

33. _____ is the energy required to break bonds.

34. _____ properties depend on the number of solute particles in the solution, not the identity of the particles.

35. The Van't Hoff factor (*i*) tells how many ions one unit of a substance will _____ into a solution.

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36. Determine the Van't Hoff factor for the following:

COMPOUND	WHAT IT DISSOCIATES INTO	VALUE OF (i)
$C_6H_{12}O_6$		
NaCl		
HNO_3		
$CaCl_2$		

37. Why does water's boiling point increase as you add salt to it?

38. Freezing point of depression

a. What causes ice to melt when salt is added to it?

b. What makes magnesium chloride a better choice to de-ice streets over sodium chloride?

39. What would increase the boiling point of water more? The same amount of sugar or table salt added to it and why?

40. Colligative properties

a. Complete the table below:

COMPOUND	WHAT IT DISSOCIATES INTO	VALUE OF (<i>i</i>)
Na ₂ S		
CaF ₂		
Al ₂ O ₃		
CCl ₄		
NaOH		

- b. Of the five compounds listed above, which one would have the highest boiling point of elevation when added to water?
- c. Of the five compounds listed above, which one would have the warmest freezing point of depression?

CHECKPOINTS

41. If solutions are produced from each of these compounds, which solutions would most likely conduct an electrical current? *Select all correct answers.*
- aluminum chloride, AlCl₃
 - chromium trioxide, CrO₃
 - ethanol, C₂H₅OH
 - glucose, C₆H₁₂O₆
 - sodium azide, NaN₃
42. Which statement correctly describes the effect of an increase in temperature on the solubility of gases dissolved in a liquid?
- Solubility increases for all gases.
 - Solubility decreases for all gases.
 - Solubility increases for most gases but decreases for some.
 - Solubility decreases for most gases but increases for some.
43. Which of the following types of compounds is most likely to be a strong electrolyte?
- a polar compound
 - a nonpolar compound
 - a covalent compound
 - an ionic compound
44. *Select the correct terms to complete the statement.* The freezing point of water when a solute is dissolved in the water is higher | lower than the freezing point of pure water. The change in the freezing point occurs because the vapor pressure of the solution is higher | lower than that of

the pure water. This change in vapor pressure also causes the boiling point of the solution to be higher | lower than the boiling point of pure water. The change in boiling point is directly proportional to the number of | identity of the solute particles.

45. A 3.25 L solution is prepared by dissolving 285 g of BaBr₂ in water. Determine the molarity. Report your final answer using the correct number of significant figures.
46. For which combination of substances does pressure have the greatest effect on the solubility?
- solids in solids
 - gases in liquids
 - solids in liquids
 - liquids in liquids
47. Which statement describes the colligative properties of a solution?
- Colligative properties are dependent on the type of solute.
 - Colligative properties do not affect the boiling point of the solution.
 - Colligative properties do not affect the freezing point of the solution.
 - Colligative properties are dependent on the concentration of the solute.
48. Which statement describes two immiscible substances?
- They will not form a solution.
 - They exist together in one phase.
 - They will not separate on standing.
 - They dissolve freely in one another in any proportion.
49. The electrical force (F) between two-point charges (q) is directly related to the product of the charges and inversely related to the distance (d) between them. What happens to an electrical force if the distance between two charges doubles while all other factors remain constant?
- $$F_{electric} = k_c \frac{q_1 q_2}{d^2}$$
- 2F
 - 4F
 - 1/2 F
 - 1/4 F
50. A solute crystal is dropped into a solution containing dissolved solute. It falls to the bottom of the beaker and does not dissolve after vigorous stirring. What does this indicate about the solution?
- It is not at equilibrium.
 - It is probably saturated.
 - It is probably unsaturated.
 - It is probably supersaturated.

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51. An apparatus for testing conductivity is placed in a solution. The power supply is turned on, and the light bulb glows brightly. Which property of the solution can be concluded from the conductivity test?
- It is saturated.
 - It is heterogeneous.
 - It is supersaturated.
 - It contains an electrolyte.
52. Which action is done by the slightly charged parts of a water molecule in a salt solution?
- They attract nonpolar molecules.
 - They attract the ions in ionic compounds.
 - They form ionic bonds with other water molecules.
 - They form ionic bonds with ions in ionic compounds.
53. Calculate the mass of Na_2CO_3 needed to make 250 mL of a 0.20 M solution of sodium carbonate.
54. Explain why solutions look like pure substances instead of mixtures with distinguishable components.
55. Explain a difference between solutes and solvents in solutions.