Name: ____ **AP Biology**
[redacted]

Chapter 2 Active Reading Guide
The Chemical Context of Life

This chapter covers the basics that you may have learned in your chemistry class. Whether your teacher goes over this chapter, or assigns it for you to review on your own, the questions that follow should help you focus on the most important points.

Section 1

1. Define and give an example of the following terms:

Term	**Definition**
matter:	
element:	
compound:	

2. What four elements make up 96% of all living matter?

3. What is the difference between an *essential element* and a *trace element*?

Section 2

4. Sketch a model of an atom of helium, showing the electrons, protons, neutrons, and atomic nucleus.

5. What is the atomic number of helium? _____ Its atomic mass? _____

6. Here are some more terms that you should firmly grasp. Define each term.

Term	**Definition**
neutron:	
proton:	
electron:	
atomic mass:	
atomic number:	
isotope:	
electron shells:	
energy:	

7. Consider the entry in the periodic table for carbon.

   What is the atomic mass? _____ What is the atomic number? _____

   How many electrons does carbon have? _____ How many neutrons? _____

8. What are *isotopes*? Use carbon as an example.

9. Explain radioactive isotopes and one medical application that uses them.

10. Which is the only subatomic particle that is directly involved in the chemical reactions between atoms?

11. What is *potential energy*?

12. Explain which has more potential energy in each pair:
    a. boy at the top of a slide/boy at the bottom

    b. electron in the first energy shell/electron in the third energy shell

    c. water/glucose

13. What determines the chemical behavior of an atom?

14. Sketch an electron distribution diagram for sodium:

    a. How many valence electrons does it have? _____
       Circle the valence electron(s).

    b. How many protons does it have? _____

Section 3

15. Define *molecule*.

16. Now, refer back to your definition of a *compound* and fill in the following chart:

	Molecule? (y/n)	Compound? (y/n)	Molecular Formula	Structural Formula
Water				
Carbon Dioxide				
Methane				
Oxygen				

17. What type of bond is seen in O_2? Explain what this means.

18. What is meant by *electronegativity*?

19. Explain the difference between a *nonpolar covalent bond* and a *polar covalent bond*.

20. Make an electron distribution diagram of water. Which element is most electronegative? Why is water considered a *polar* molecule? Label the regions that are more positive or more negative. (This is a very important concept. Spend some time with this one!)

21. Another bond type is the *ionic bond*. Explain what is happening in Figure 2.10.

22. What two elements are involved above?

23. Define *anion* and *cation*. In the preceding example, which is the anion?

24. What is a *hydrogen bond*? Indicate where the hydrogen bond occurs in Figure 2.12.

25. Explain *van der Waals interactions*. Though they represent very weak attractions, when these interactions are numerous they can stick a gecko to the ceiling!

26. Here is a list of the types of bonds and interactions discussed in this section. Place them in order from the strongest to the weakest: hydrogen bonds, covalent bonds, ionic bonds, van der Waals interactions.

STRONG

WEAK

27. Use morphine and endorphins as examples to explain why molecular shape is crucial in biology.

Section 4

28. Write the chemical shorthand equation for photosynthesis. Label the *reactants* and the *products*.

29. For the equation you just wrote,
    How many molecules of carbon dioxide are there? _____

    How many molecules of glucose? _____

    How many elements in glucose? _____

30. What is meant by *dynamic equilibrium*? Does this imply equal concentrations of each reactant and product?

Section 5

31. What is a *polar molecule*? Why is water considered polar?

32. Explain *hydrogen bonding*. How many hydrogen bonds can a single water molecule form?

33. Distinguish between *cohesion* and *adhesion*.

34. Which is demonstrated when you see beads of water on a waxed car hood?

35. Which property explains the ability of a water strider to walk on water?

36. The calorie is a unit of heat. Define *calorie*.

37. Water has high *specific heat*. What does this mean? How does water's specific heat compare to alcohol's specific heat?

38. Explain how hydrogen bonding contributes to water's high specific heat.

39. Summarize how water's high specific heat contributes to the moderation of temperature. How is this property important to life?

40. Define *evaporation*. What is *heat of vaporization*? Explain at least three effects of this property on living organisms.

41. Ice floats! So what? Consider what would happen if ponds and other bodies of water accumulated ice at the bottom. Describe why this property of water is important.

42. Now, explain *why* ice floats. Why is 4°C the critical temperature?

43. Review and define these terms:

Term	**Definition**
solvent:	
solution:	
solute:	

44. Consider coffee to which you have added sugar. Which of these is the solvent? Which is the solute?

45. Explain why water is such a fine solvent.

46. Distinguish between *hydrophobic* and *hydrophilic substances*. Give an example of each.

47. You already know that some materials, such as olive oil, will not dissolve in water. In fact, oil will float on top of water. Explain this property in terms of hydrogen bonding.

48. Now, let's do a little work that will enable you to prepare solutions. Read the section on solute concentrations carefully, and show the calculations here for preparing a 1-molar solution of sucrose. Steps to help you do this follow. The first step is done for you. Fill in the rest.

Steps to prepare a solution:

a. Write the molecular formula.

$C_{12}H_{22}O_{11}$

b. Use the periodic table (on Page B-1) to calculate the mass of each element. Multiply by the number of atoms of the element. (For example, O has a mass of 16. Therefore, one mole of O has a mass of 16 x 11 = 176 g/mole.)

c. Add the masses of each element in the molecule.

d. Add this mass of the compound to water to bring it to a volume of 1 liter. This makes 1 liter of a 1 *M* (1-molar) solution.

49. Can you prepare 1 liter of a 0.5-molar *glucose* solution? Show your work here.

50. Define *molarity.*

51. What two ions form when water dissociates?

52. What is the concentration of each ion in pure water at 25°C?

53. *pH* is defined as the negative log of the hydrogen ion concentration $[H^+]$. Explain how water is assigned a pH of 7.

54. To go a step further, the product of H^+ and $OH^–$ concentrations is constant at $10^{–14}$.

$$[H^+]\,[OH^–] = 10^{–14}$$

Water, which is neutral with a pH of 7, has an equal number of H^+ and $OH^–$ ions. Now, define

Acid:

Base:

55. Because the pH scale is logarithmic, each numerical change represents a 10X change in ion concentration.
    a. How many times more acidic is a pH of 3 compared to a pH of 5? _____
    b. How many times more basic is a pH of 12 compared to a pH of 8? _____
    c. Explain the difference between a pH of 8 and a pH of 12 in terms of H^+ concentration.

56. Even a slight change in pH can be harmful! How do *buffers* moderate pH change?

57. Exercise will result in the production of CO_2, which will acidify the blood. Explain the buffering system that minimizes blood pH changes.

58. *Acid precipitation* is increasing. What is the pH of uncontaminated rain?

59. Give two reasons precipitation is more acidic today compared to 1900.

60. What products of fossil fuel burning contribute to acid precipitation?

61. Discuss how CO_2 emissions affect marine life and ecosystems.