**CH301 Worksheet 13b Answer Key—Internal Energy Lecture**

1.  Define  open, closed, or isolated systems.   If you use an open system as a calorimeter, what is the state function you can calculate from the temperature change.  If you use a closed system as a calorimeter, what is the state function you can calculate from the temperature?

**Answer:**   An **open system** can exchange both matter and energy with the surroundings. $\Delta$ H is measured when an open system is used as a calorimeter.

 A **closed system** has a fixed amount of matter, but it can exchange energy with the surroundings. $\Delta$ U is measured when a closed system is used as a calorimeter because there is no change in volume and thus no expansion work can be done.

An **isolated system** has no contact with its surroundings. The **universe** is considered an isolated system but on a less profound scale, your thermos for keeping liquids hot approximates an isolated system.

2.  Rank, from greatest to least, the types internal energy found in a chemical system:

**Answer:**   The energy that holds the nucleus together  is much greater than the energy in chemical bonds (covalent, metallic, network, ionic) which is much greater than IMF (Hydrogen bonding, dipole, London)  which depending on temperature are approximate in value to  motional energy (vibrational, rotational, translational).

3.  Internal energy is a state function.  Work and heat (w and q) are not.  Explain.

**Answer:**   A **state function (like U, V, T, S, G)** depends only on the current state of the system so if the system is changed from one state to another, the change in a state function is independent of the path.  When we calculate $\Delta$ of U, V, T, S, G we are indicating that we are only interest in how things ended compared to how things began.

In contrast, **work and heat (w and q)** are not state functions and instead represent the amounts of energy recorded **along the path** between changes in state.

4.  What are the two types of work most commonly associated with chemical processes?

**Answer:  Expansion work** which requires changing the volume of a system, typically a gas, and **electrical work** which requires creation of electrical current from the controlled movement of electrons  in  an oxidation-reduction reaction.

5.  Calculation from the text on the most famous kind of expansion work NOT done by a gas:

Water expands when it freezes. How much work does 100  g of water do when it freezes at 0°C and bursts a water pipe that exerts an opposing pressure of 1070 atm? The densities of water and ice at 0°C are 1.00 g·cm$^{-3}$ and 0.92 g·cm$^{-3}$, respectively.
**Answer:**  w = −0.9 kJ

6.  Name the two phase changes for which significant work is done on the system?
**Answer:**  Condensation (conversion of a gas into a liquid)  and deposition  (conversion of a gas into a solid.)

7.  Distinguish  heat capacity, specific heat and molar heat capacity and provide typical examples of units?
**Answer:**  Heat capacity is the amount of heat it takes to raise the temperature of a substance a certain amount. The equation used is $C = q/T$ $(J \cdot K^{-1})$ while the other two types of heat capacity are per specific amounts of substance.  Specific heat is $C_s = q/T$ for a gram of substance $(J \cdot K^{-1} \cdot g^{-1})$ and molar heat capacity $C_m = q/T$ for a mole of substance $(J \cdot K^{-1} \cdot mole^{-1})$

8. Rank the following in terms of increasing heat capacity. Water, ice, copper, air.
**Answer:** Based on specific heat (per gram) water > ice > air > copper

9. A calorimetry calculation from the text:

Potassium perchlorate, $KClO_4$, is used as an oxidizer in fireworks. Calculate the heat required to raise the temperature of 10.0 g of $KClO_4$ from 25°C to an ignition temperature of 900.°C. The specific heat capacity of $KClO_4$ is 0.8111 $J·K^{-1}·g^{-1}$.
**Answer:** 7.10 kJ

10. The change is internal energy is found by measuring the amount of heat or work that enters or leaves the system. Use this fact to perform the following calculation from the text:

An automobile engine does 520. kJ of work and loses 220. kJ of energy as heat. What is the change in the internal energy of the engine? Treat the engine, fuel, and exhaust gases as a closed system.

**Answer:** −740. kJ

11. Were it not for the fact that most thermochemistry in real life is done in open systems at constant external atmospheric pressure, we would be happy to use $\Delta U$ to account for energy changes in a system. But in the real world, $\Delta H$ is the more practical unit of energy change in a system. Given the definition of $H = U + PV$, derive the fact that a calorimeter in at open system at constant external pressure has $\Delta H = q_p$.

**Answer:** Consider a process at constant pressure for which the change in internal energy is $\Delta U$ and the change in volume is $\Delta V$. It then follows from the definition of enthalpy that the change in enthalpy is

$$\text{At constant pressure: } \Delta H = \Delta U + P\Delta V$$

Now we use $\Delta U = q + w$, where q is the energy supplied to the system as heat and w is the energy supplied as work. Therefore,

$$\Delta H = q + w + P\Delta V$$

Next, suppose that the system can do no work other than expansion work. In that case, $w = -P_{ex}\Delta V$ so

$$\Delta H = q - P_{ex}\Delta V + P\Delta V$$

Finally, because the system is open to the atmosphere, the pressure of the system is the same as the external pressure; so $P_{ex} = P$, and the last two terms cancel to leave $\Delta H = q$.

12. Apply what you know about the relationship between H and U to perform the following calculation from the text:

In an exothermic reaction at constant pressure, 50. kJ of energy left the system as heat and 20. kJ of energy left the system as expansion work. What are the values of (a) $\Delta H$ and (b) $\Delta U$ for this process?
**Answer:** (a) −50. kJ; (b) −70. kJ